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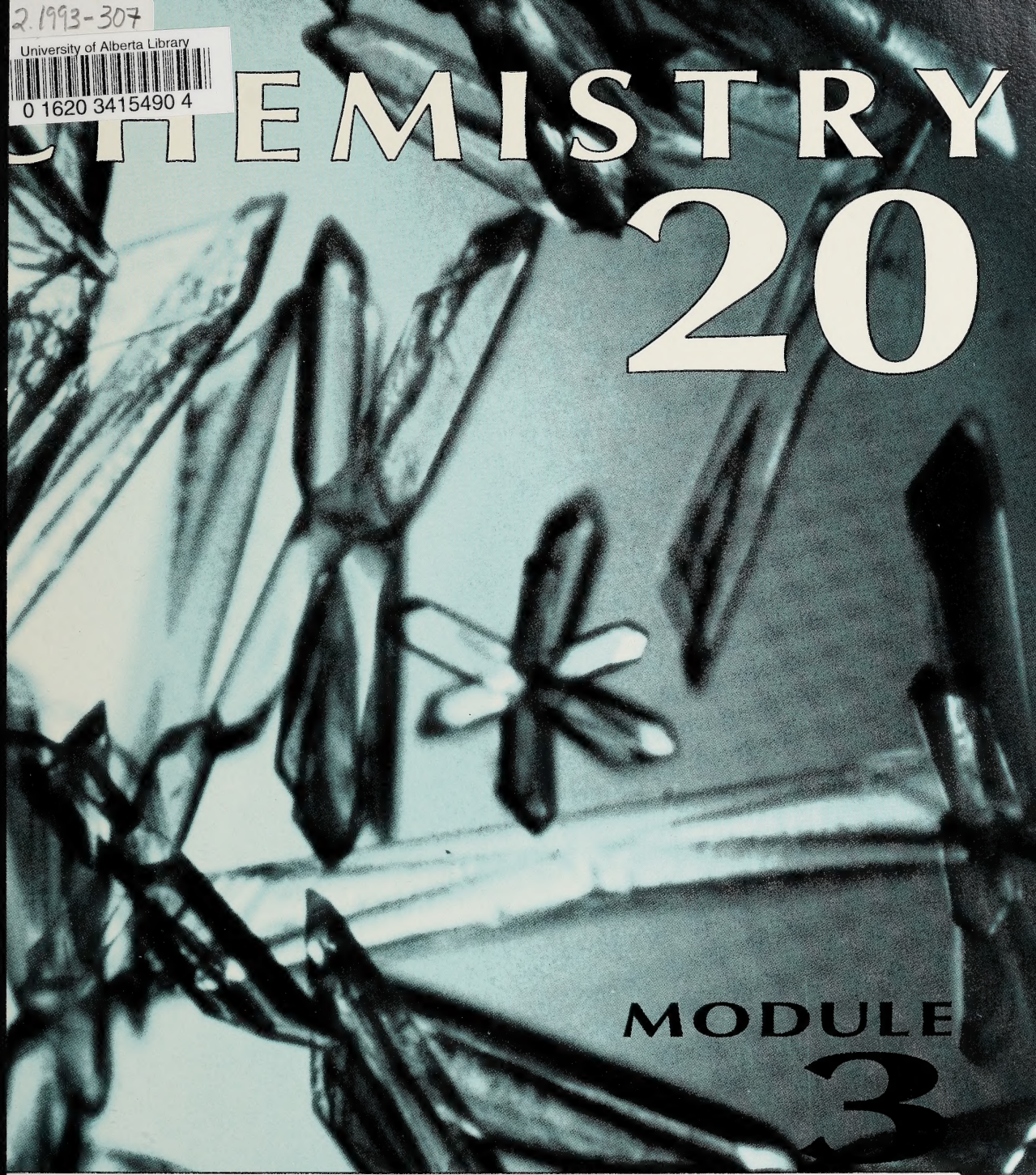
CHEMISTRY 20



MODULE 3

QUANTITATIVE RELATIONSHIPS

Distance Learning

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Module 3

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Chemistry 20
Student Module
Module 3
Quantitative Relationships
Alberta Distance Learning Centre
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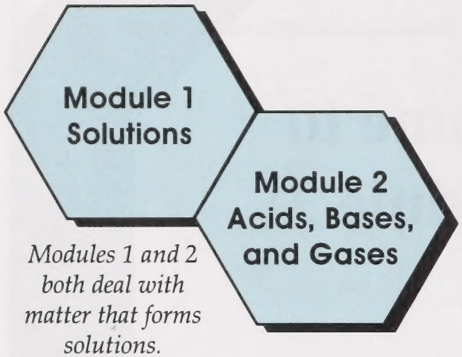
Welcome to Module 3!

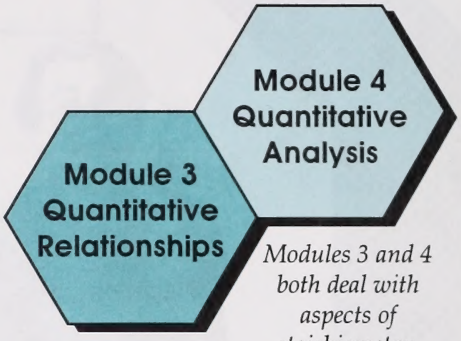
We hope you'll enjoy your study of *Quantitative Relationships*. Remember, to make your learning a bit easier, several videocassettes and laser videodiscs are referred to throughout the modules. Your textbook *Nelson Chemistry* will also enhance your studies.

COURSE OVERVIEW

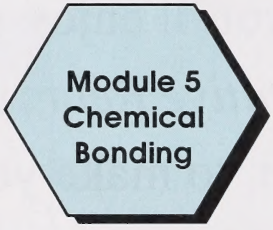
This course contains seven modules. The module you are working in is highlighted in deeper colour.

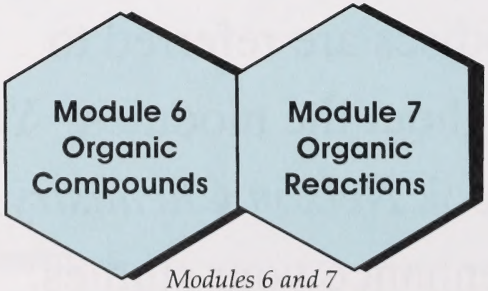


Modules 1 and 2 both deal with matter that forms solutions.



Modules 3 and 4 both deal with aspects of stoichiometry.





Modules 6 and 7 both deal with organic chemistry.

MODULE 3

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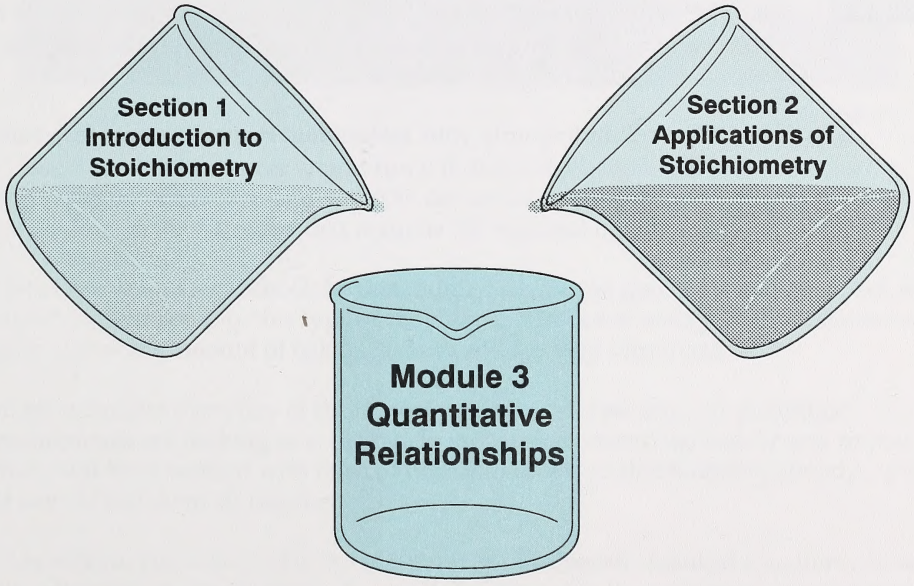
MODULE OVERVIEW

Is it important for you to know how many ASA tablets to take if you have a headache? Would it matter how much salt you added to the batter when making your favourite cookies? Everyone can appreciate the need for appropriate quantities and proportions in dealing with many aspects of everyday life. Doctors are extremely careful in prescribing the correct dosage of medicine for their patients. Bakers know the importance of having the correct amount of flour in the bread that they bake. You even make sure that you have the right proportion of water and lemon when making lemonade.

You can see, then, the need for chemists to know the correct proportions of chemicals needed to cause a complete reaction. Chemistry in industry must also consider the amounts of chemicals in reactions so as not to cause waste.

In this module you will examine the role that quantities of matter play in chemical reactions. You will get a chance to use knowledge you already have about chemical reactions and apply this knowledge to industrial and everyday situations.



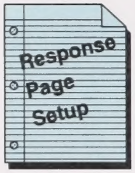
Evaluation

Your mark in this module will be determined by how well you complete the assignments at the end of each section. You must complete all assignments. In this module you are expected to complete two assignments. The mark distribution is as follows:

Section 1 Assignment	**50 marks**
Section 2 Assignment	**50 marks**
TOTAL	**100 marks**

When doing your assignments, work slowly and carefully. If you are having difficulties, go back and review the appropriate section.

Read all parts of your assignment carefully. Plan and do your rough work on your own paper. Revise and edit your responses; then set up your final copy for submission on your own paper. Lined looseleaf is recommended. Make sure your answers are neat and organized, with wide left margins and space for teacher comments after each assignment.



When you see this icon, ideas and details are provided to help you set up and organize your answer in a certain way.

Before submitting your responses, be sure to proofread them carefully to ensure they say what you want, that they are neat and clear, and that they are complete and missing no material.

You will be submitting **only** your **assignment response pages** for evaluation.

SECTION 1

Introduction to Stoichiometry

WESTFILE INC.

How well do you think a car would run if it didn't have some method of mixing the correct amount of gasoline and air? The car would likely not run at all. All vehicles have devices on their engines that regulate the required quantities of air and fuel.

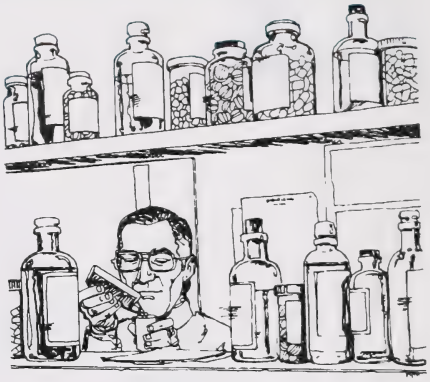
Of what use would a cookbook be that didn't describe the quantities of ingredients in its recipes? Would you buy this type of cookbook? Directions telling you the number of eggs to use or the amount of baking soda to add are very important.

You see examples everyday of the importance of measurements and quantities. Measurements are nothing new to you. In fact, stoichiometry isn't really new to you, either. You have worked with most of the components of stoichiometry already – you just have to put them all together.

In this section, you will review the components of balanced chemical equations, mole to mass problems, and mole ratios. You will discover what the term stoichiometry means. As well, you will get a chance to combine the components you have worked with into a full stoichiometric process.

ACTIVITY 1

Equations and Mole Ratios



By the time you have finished this chemistry course, you will have worked with many different chemical reactions.

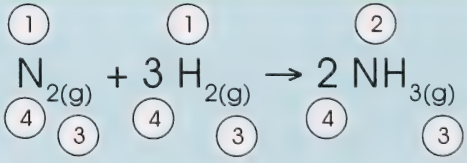
Professional chemists work with hundreds of different reactions, some in the lab, but many on paper.

1. How can chemical reactions be represented on paper?

chemical equation – the formulas of the reactants on the left and the formulas of the products represented on the right connected with an arrow representing a chemical reaction

You learned in Science 10 how to balance **chemical equations**. Chemical equations show how chemical substances react to form products. They also illustrate the proportions of chemicals involved. You also had some practice with writing equations in Module 1. This activity will give you more opportunities to write and balance chemical equations, as well as work with the proportions of chemicals in terms of mole ratios.

What are the important factors expressed in a balanced chemical equation?



1. Chemical formulas of reactants
2. Chemical formula of product
3. States of substances – can be s, l, g, or aq
4. Balancing coefficients

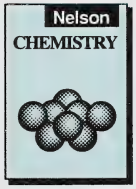
Balanced chemical equations give you a lot of information and are useful in predicting expected products of reactions.

You can review the steps required to balance a chemical equation by reading pages 98 to 100 in your textbook.

chemical formulas – a shorthand way of showing the number and type of atoms present in a substance

2. Why do you think that the most important step is to have the correct **chemical formulas** of the reactants and products?

3. Complete questions 2 to 6 from page 100 of your textbook. Note: The perspectives (points of view) that are referred to are described on pages 42 and 43 of your textbook.

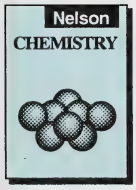


Check your answers by turning to the Appendix, Section 1: Activity 1.

Chemical reactions are usually classified on the basis of what reacts or what is produced. Review what you have previously learned about reaction types by reading Section 4.4: Classifying Chemical Reactions on pages 101-103 of your textbook.

4. Complete the following table on chemical reaction types.

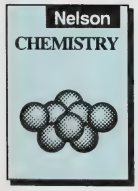
Reaction Type	General Description
formation	elements → _____
_____ decomposition	_____ → elements
_____	substance + oxygen → _____
_____	_____ + _____ → element + compound
double replacement	_____ + _____ → _____ + _____



5. Use a similar chart for completing question 7 from page 103 of your textbook.

Textbook Question	Reaction Type	Balanced Equation
a		
b		

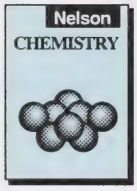
Check your answers by turning to the Appendix, Section 1: Activity 1.



You probably realized that only three types of reactions were balanced in the previous question. The other two general types of reactions can be reviewed by reading pages 103-106 in your textbook. This section also reviews key aspects of solution chemistry that were discussed in Module 1.

6. Match the following terms with their definitions.

Terms	Definitions
a. solubility	i. solid substances formed as a result of an aqueous reaction
b. precipitate	ii. the substance doing or causing the dissolving
c. solute	iii. dissolves poorly in solution
d. high solubility	iv. the substance that dissolves in solution
e. low solubility	v. the degree to which a substance dissolves in solution
f. solvent	vi. dissolves well in solution



7. Using the solubility chart on the back cover of your textbook, determine the solubility of each of the following substances. Write the chemical formula and appropriate state for each substance.

   a. sodium sulphate

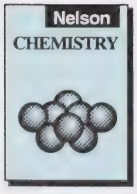
   b. copper(II) bromide

   c. calcium carbonate

   d. potassium sulphite

   e. aluminum hydroxide

8. Why did your textbook separate single and double replacement reactions from the other reaction types studied previously?

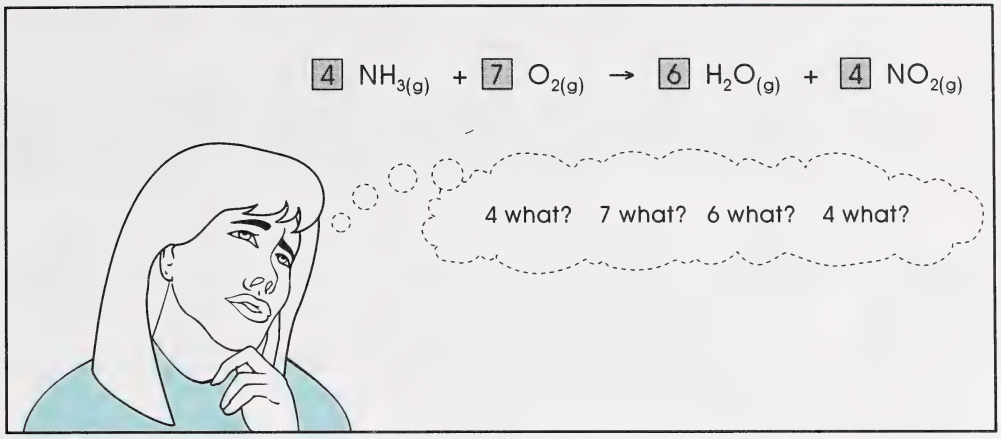


9. Complete question 14 (a, b, c, d, e, f) from page 109 of your textbook.

10. Complete question 8 (a, b) from page 107 of your textbook.

Check your answers by turning to the Appendix, Section 1: Activity 1.

You have already seen the need for coefficients in chemical equations – they balance the equation so the number of atoms of each element is equal on both sides of the equation. But what information do the coefficients give you with respect to the actual chemical reaction? Do the coefficients tell you practical amounts of chemicals to react in the lab?



The preceding questions will be answered in the videos from *The Mole Concept Series*. You may have already seen three videos in this series in Module 2.

View the video Relative Mass from *The Mole Concept Series*. This video will give you an idea about what amount is practical when dealing with chemical substances. Preview the following questions before watching the video.

11. Why is it important to be able to compare the relative masses of atoms?

relative mass – the mass of one substance compared to another when equal numbers of both substances are used

12. Why do chemists use the term **relative mass** when talking about atoms?

13. Why did the narrator speak of finding a universal container?

14. When comparing two different substances, do you have to use just one unit of each to compare relative mass? Explain.

Check your answers by turning to the Appendix, Section 1: Activity 1.

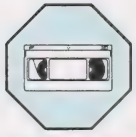
The term relative mass is useful in some ways for comparing two substances, but to actually measure an amount of a substance, you need something more practical. View the video Relative Atomic Mass from *The Mole Concept Series* to see how relative mass can be made more practical. Preview the following questions before watching the video.

15. a. What element was first used as a standard to find the relative mass of other molecules?

    b. What element provides the standard mass accepted today?

16. Today, what instrument allows scientists to find the relative mass of atoms of any substance, not just molecules of gas?

Check your answers by turning to the Appendix, Section 1: Activity 1.



The question still remains: How does the relative atomic mass of atoms relate to the number of atoms participating in a chemical reaction? Also, is a universal container something that is possible? Hopefully, the answers to these questions will become clear when you watch the video The Mole from *The Mole Concept Series*. Preview the following questions before watching the video.

relative atomic mass – the mass of one substance compared to another (standard – Carbon-12) when equal numbers of atoms are present

17. What is the problem in working with volumes and a universal container for gases only?

18. Why wouldn't an Aardvark work as a universal container?

19. What did scientists pick as the useful number of atoms?

20. How many atoms are in 12 grams of Carbon-12?

21. How can you use the relative mass of an element to determine the number of atoms of that element you have?

22. What is this universal container of atoms called?

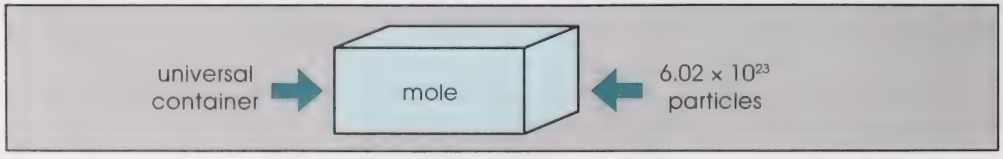
mole – the amount of substance that contains 6.02×10^{23} particles

23. Under what conditions do you have one mole of a gas?

24. Approximately how many atoms are there in a mole?

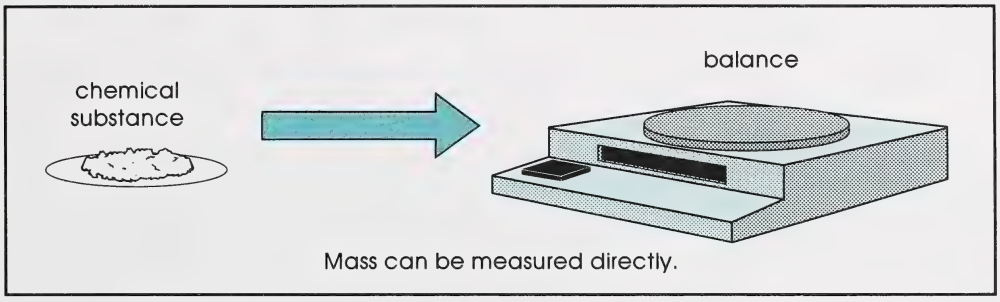
25. Why was it so important for chemists to develop the concept of the mole?

Check your answers by turning to the Appendix, Section 1: Activity 1.



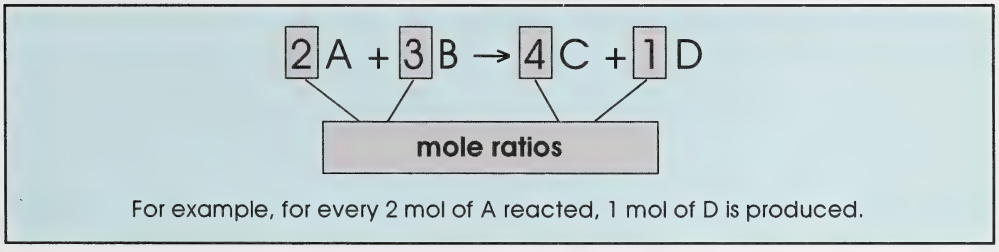
From the videos you learned that it is not possible to measure the mass of single atoms for comparison. You also saw that early chemists were able to compare ratios of the masses of different gases in such a way as to make meaningful comparisons. Modern scientists using sophisticated technology have compared the mass of all the different types of atoms. This knowledge of relative mass has many important applications in chemistry.

The connection between mass and moles, as you discovered in the videos, is very important in dealing with chemical reactions. The mass unit provides a practical means of measuring materials in the lab.



Mass can be measured directly.

The mole unit allows relative amounts of chemicals to be compared in chemical equations.



For example, for every 2 mol of A reacted, 1 mol of D is produced.

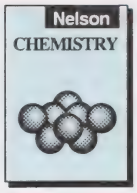
Nelson CHEMISTRY

The universal container, the mole, enables you to make a connection between the amount of chemical required, as shown by the balanced equation, and the practical mass of the chemical that can be measured. To learn more about moles and chemical equations, read the bottom half of page 97 and the top two-thirds of page 98 in your textbook.

Avogadro's constant – the number of particles (6.02×10^{23}) in one mole of a substance

26. What is the SI symbol for Avogadro's constant (Avogadro's number)?

27. Why can you read a balanced chemical equation in terms of individual molecules, large numbers of molecules, or moles?



28. Complete question 1 (a, b) from page 98 of your textbook.

Check your answers by turning to the Appendix, Section 1: Activity 1.

In this activity, you reviewed some important concepts including how to compare atoms of different substances by using the quantity known as a mole. You also refreshed your memory on balancing equations.

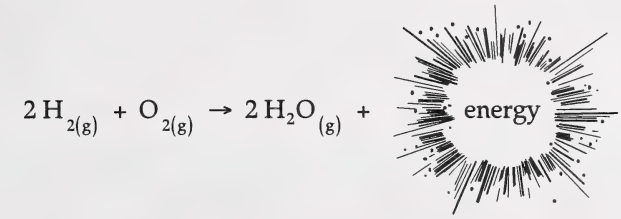
Knowledge of moles and equations allows chemists in industry to know how much of one substance can be added to another substance to produce a product in economical amounts.

ACTIVITY 2

What Is Stoichiometry?

NASA

Did you realize that by making water you could produce enough energy to make a rocket blast off? Liquid hydrogen is one of the fuels that powers rockets such as the space shuttle. When hydrogen vapourizes and burns, it produces a tremendous amount of energy.



What information does the balanced equation give you? One piece of information the balanced equation describes is the proportions in which substances react.

$$2\,H_{2(g)} + O_{2(g)} \rightarrow 2\,H_2O_{(g)}$$

PROPORTIONS

2 parts	1 part	2 parts
2 molecules	1 molecule	2 molecules
2 mol	1 mol	2 mol

mole ratio – a whole number ratio between two substances involved in a chemical reaction

stoichiometry – the use of proportions to calculate quantities of substances involved in chemical reactions

Another term for these proportions is **mole ratio**. You can use this information to calculate the amounts of substances involved in chemical reactions. The use of proportions to calculate quantities of chemicals reacted and produced using chemical equations is called **stoichiometry**.

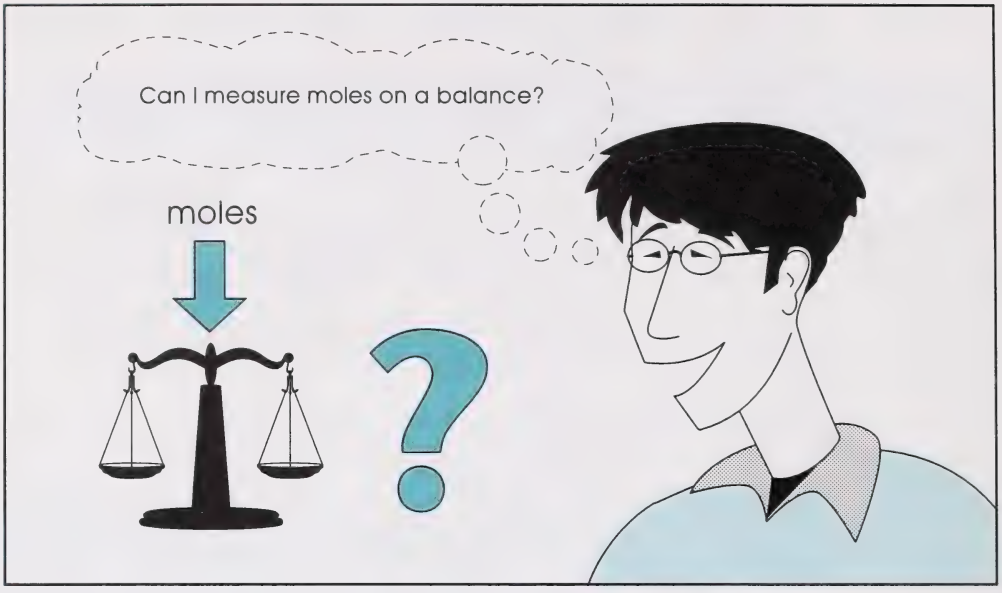
Rocket scientists use stoichiometry to make sure that they have the fuel mixed in the right proportions with oxygen.

1. If the huge external tank of the shuttle spacecraft holds 2×10^7 mol of hydrogen, how much liquid hydrogen in moles should the tank hold to get a complete reaction?

Check your answers by turning to the Appendix, Section 1: Activity 2.

The external tank of the shuttle actually contains 5×10^7 mol of hydrogen, which means there is more hydrogen than is needed (hydrogen is in excess) to react with the oxygen.

You will find that supplying slightly more of one reactant than is needed is quite common. This is done to assure that the other reactant is completely reacted.



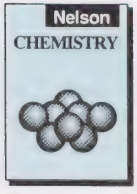
molar mass – the number of grams of a chemical substance in one mole (6.02×10^{23} particles) of that substance

Factor Label method – a method of solving all the steps of a problem by using units arranged on one line so that the units can be cancelled, revealing the final unit of the solution

The unit mole refers to an amount or number of objects or particles, not to a unit like grams which can be measured on a balance. You use mole units when dealing with chemicals in balanced equations, but you use mass units to actually measure the chemicals. Therefore, you need a way to convert between moles and mass. In Science 10, you learned that this conversion is achieved by using the molar mass of the chemical.

You may feel you need a review of molar mass and mass-amount (moles) conversions. If so, read the top 3/4 of page 156 in your textbook and answer the following questions.

Note: Your textbook usually uses the Factor Label (conversion factor) method of solving problems, as you used in Module 1. However, you can use the formula method for solving many of the problems, as you also learned in Module 1 and as is mentioned on page 156 of your textbook.



2. Complete question 15 (a, b, c, d, e, f, g), 16, 17 (a, b, c), and 18 (a, b, c) from page 157 of your textbook. Remember to consider the significant digits in your answers.

Check your answers by turning to the Appendix, Section 1: Activity 2.

Earlier in this activity, you read the discussion about hydrogen as a fuel on the shuttle spacecraft. You were given the amount in moles of hydrogen stored in the external tank. If a mole is not a measurable unit, how do you think technicians measure the actual amount of hydrogen and oxygen needed?

Since the hydrogen and oxygen are in liquid form, the fuel is measured either in litres or by mass (kilograms). For example, how many kilograms of liquid hydrogen would contain the 5.0×10^7 moles of hydrogen needed? To solve this problem, you have to convert the known moles to the required mass.

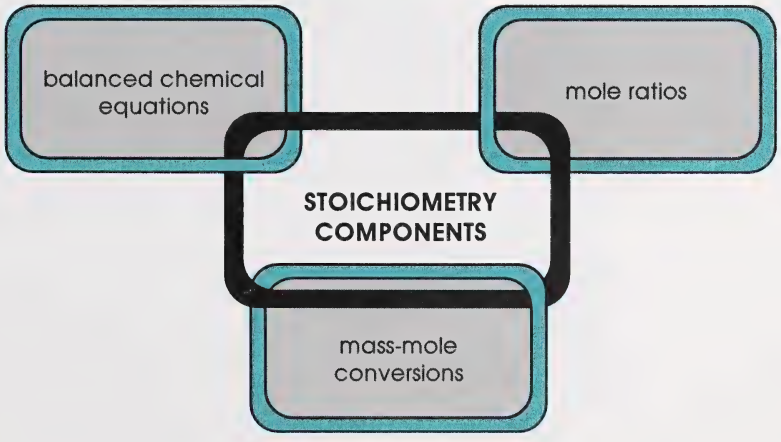
List of Variables: $m = ?$

$n = 5.0 \times 10^7$ mol H_2

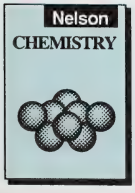
$M = 2.02$ g/mol H_2 (molar mass)

Formula Method	Factor Label Method
$n = \frac{m}{M}$ $m = nM$ $= 5.0 \times 10^7 \text{ mol } H_2 \times \frac{2.02 \text{ g}}{1 \text{ mol } H_2}$ $= 1.0 \times 10^8 \text{ g } H_2$ $= 1.0 \times 10^5 \text{ kg } H_2$	$m_{(H_2)} = 5.0 \times 10^7 \cancel{\text{mol H}_2} \times \frac{2.02 \text{ g}}{1 \cancel{\text{mol H}_2}}$ $= 1.0 \times 10^8$ g $= 1.0 \times 10^5$ kg The molar mass was used as a conversion factor.

You have now seen and worked with all the components of one type of stoichiometry problem.



gravimetric stoichiometry – stoichiometry dealing with mass and mole quantities



The process which involves calculating the masses of reactants and products in a chemical reaction is called gravimetric stoichiometry. To learn more about gravimetric stoichiometry, read pages 169-171 in your textbook.

3. What do the coefficients of a balanced chemical equation indicate?

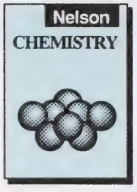
4. What does the molar mass allow you to do?

5. In the two example stoichiometry problems given in your textbook on pages 170 and 171, where was the given information (measurements) written to keep track of which information goes with which substance?

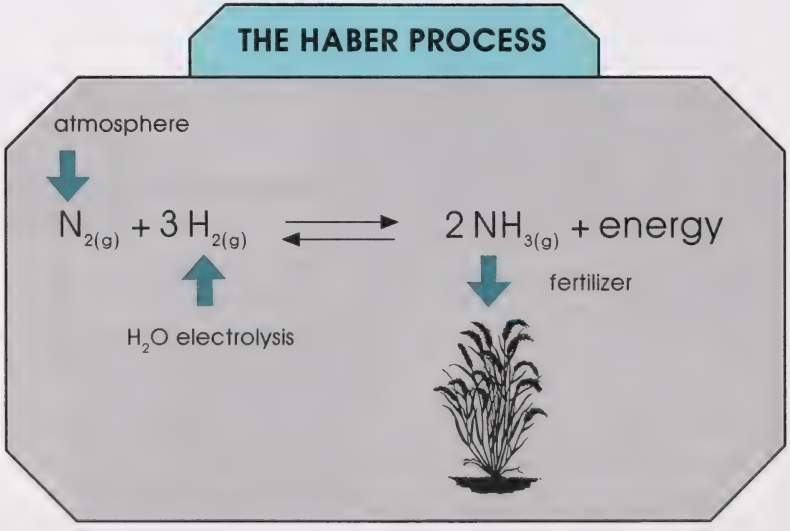
Check your answers by turning to the Appendix, Section 1: Activity 2.

Your textbook illustrated two examples of gravimetric stoichiometry problems. A third example will give you an insight into a process used in making fertilizer – the Haber process.

The Haber process (also called the Haber-Bosch process) involves atmospheric nitrogen gas reacting with hydrogen gas under controlled conditions of temperature and pressure. Ammonia, a type of fertilizer, is produced. The Haber process is a good example of utilizing principles of equilibrium as you learned in Module 1. You can also use the process to study stoichiometry principles.



Read the case study dealing with the Haber process on page 443 of your textbook.



Put yourself in the place of an industrialist using the Haber process to produce fertilizer. A customer approaches you requesting 1.000 kg (1000 g) of ammonia (the small mass used here is only for the sake of example). The question you then ask yourself is, "How many grams of nitrogen gas are needed to make 1000 g (assume four significant digits) of ammonia?"

To answer your question, follow the stoichiometry steps you have learned.

Method 1

Step 1 Write a balanced chemical equation.

$$\begin{array}{lll} N_{2(g)} + 3\,H_{2(g)} & \rightarrow & 2\,NH_{3(g)} \\ m = ? & & 1000\ g \\ 28.02\ g/mol & & 17.04\ g/mol \end{array}$$

- Underneath the equation you can write the given information (mass of $NH_{3(g)}$), the required information (mass of $N_{2(g)}$), and any other values you will need (molar masses, in this case).

Step 2 Convert the mass of ammonia to moles of ammonia.

$$n_{(NH_3)} = 1000\ \cancel{g\ NH_3} \times \frac{1\ mol}{17.04\ \cancel{g\ NH_3}}$$
$$= 58.69\ mol$$

- The molar mass factor was inverted $\left(\frac{mol}{g}\right)$ so that the grams would cancel.
- The answer is rounded to 4 significant digits (sig. dig.) because of the 4 sig. dig. in each number used in the calculation.

Step 3 Convert the moles of ammonia to moles of nitrogen using the mole ratio from the equation.

$$n_{(N_2)} = 58.69\ \cancel{mol\ NH_3} \times \frac{1\ mol\ N_2}{2\ \cancel{mol\ NH_3}}$$
$$= 29.35\ mol\ N_2$$

- The mole ratio is of the form, $\frac{x\ mol\ required}{y\ mol\ given}$, where x and y are the coefficients from the balanced equation.
- The mol units are used in the ratio as a check to see that the ratio is used the correct way – the given units $(mol\ NH_3)$ cancel.

Step 4 Convert the moles of N_2 into the mass of N_2.

$$m_{(N_2)} = 29.35 \cancel{\text{ mol } N_2} \times \frac{28.02 \text{ g}}{1 \cancel{\text{ mol } N_2}}$$
$$= 822.2 \text{ g}$$

- This means that 822.2 g of nitrogen gas, along with enough hydrogen gas for a complete reaction, will produce 1000 g of ammonia.

Method 2

You can also use formulas to do this calculation.

Step 1 $N_{2(g)} + 3\,H_{2(g)} \rightarrow 2\,NH_{3(g)}$

Step 2
$$n_{(NH_3)} = \frac{m}{M}$$
$$= \frac{1000 \text{ g}}{17.04 \text{ g/mol}}$$
$$= 58.69 \text{ mol } NH_3$$

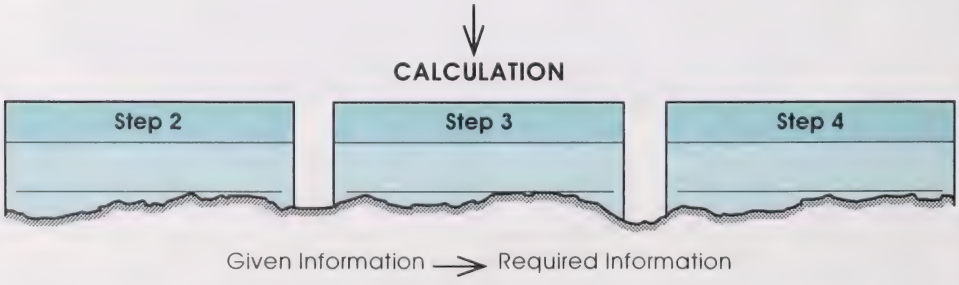
Step 3
$$n_{(N_2)} = 58.69 \text{ mol } NH_3 \times \frac{1 \text{ mol } N_2}{2 \text{ mol } NH_3}$$
$$= 29.35 \text{ mol } N_2$$

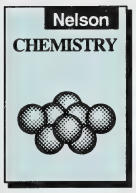
Step 4
$$m_{(N_2)} = n \times M$$
$$= 29.35 \text{ mol } N_2 \times \frac{28.02 \text{ g}}{1 \text{ mol } N_2}$$
$$= 822.2 \text{ g } N_2$$

You are free to choose the method you prefer.

6. In a chart such as the following, summarize (in words) the steps used to find the mass of a required substance from the mass of a given substance in a gravimetric stoichiometry problem.

Step 1: _____





7. Complete question 9 from page 171 of your textbook.

Check your answers by turning to the Appendix, Section 1: Activity 2.

The Factor Label method of doing the previous stoichiometry problems can be shortened further. Instead of writing down rounded answers at the end of each step and re-entering those numbers to continue on, you can set the whole problem up initially in one line.

Using the ammonia example shown previously, the following is how the setup would look when arranged in one line.

Step 1

$$N_{2(g)} + 3\,H_{2(g)} \rightarrow 2\,NH_{3(g)}$$

m = ? 1000 g

28.02 g/mol 17.04 g/mol

Step 2 Step 3 Step 4

$$m_{(N_2)} = 1000\ \cancel{g\ NH_3} \times \frac{1\ \cancel{mol\ NH_3}}{17.04\ \cancel{g\ NH_3}} \times \frac{1\ \cancel{mol\ N_2}}{2\ \cancel{mol\ NH_3}} \times \frac{28.02\ g}{1\ \cancel{mol\ N_2}}$$

$$= 822.2\ g$$

If you are having trouble using your calculator in solving the preceding type of problem, there are a number of ways the values can be entered for finding the final answer. Two possible ways are as follows.

Method 1

Enter	Display
1000	1000
×	1000
28.02	28.02
÷	28020
17.04	17.04
÷	1644.366197
2	2
=	822.1830986

All the top terms are multiplied, then all the bottom terms divided.

Method 2

Enter	Display
1000	1000
÷	1000
17.04	17.04
÷	58.68544601
2	2
×	29.342723
28.02	28.02
=	822.1830986

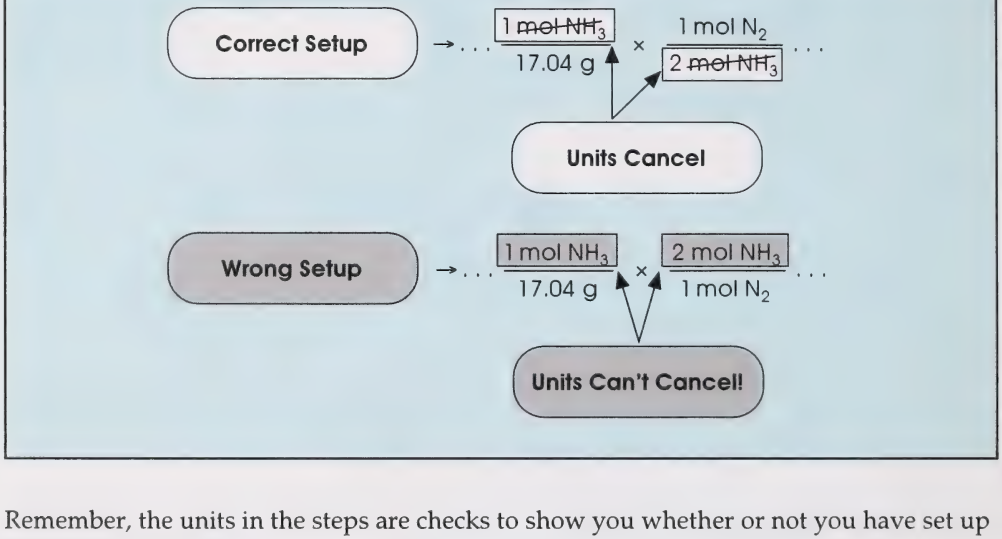
Each operation is done in the order of the steps.

Multiplying and dividing by one does not affect the answer, so it is omitted in the methods.

There are two advantages to setting up the problem in one line.

- You can set up the whole problem, without having to use your calculator for any intermediate steps.
- You only have to round off once – at the end when you have your answer.

In the one-line method, each step is set up properly by keeping track of the units. In each new step, the units in the denominator should cancel with the units in the numerator of the previous step.

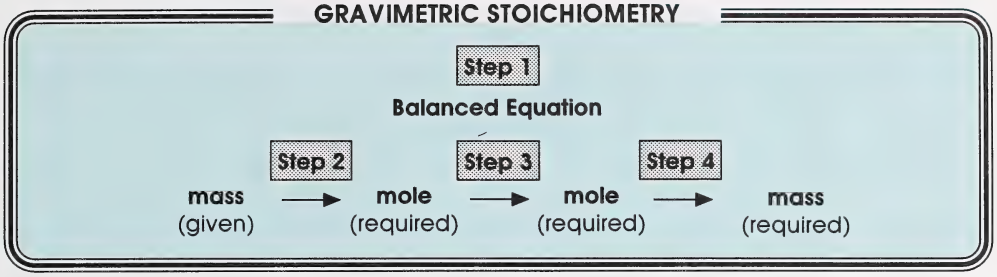


Remember, the units in the steps are checks to show you whether or not you have set up the solution properly.

It will be your choice as to the method you choose to set up your problem. The answers that will be given in the Appendix for stoichiometry problems, though, will be set up in one line.

8. To give you more practice in solving stoichiometry problems, complete questions 10 to 14 from page 171 of your textbook. (Note: Propane is $C_3H_{8(g)}$.)

Check your answers by turning to the Appendix, Section 1: Activity 2.

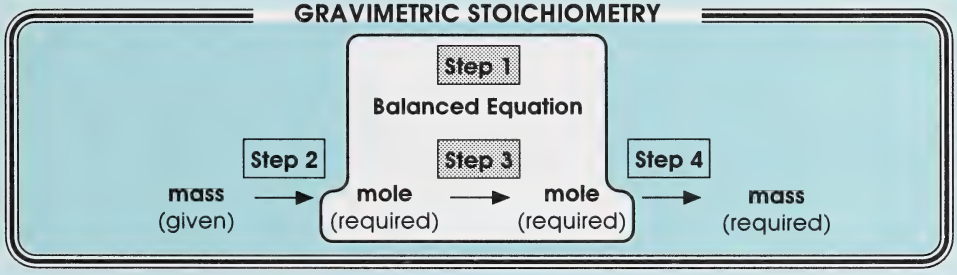


There will be times when you will solve problems that do not necessarily need all four steps of the stoichiometry process.

Example 1

How many moles of aluminum metal will react in solution with excess copper(II) chloride to produce 1.65 mol of copper metal?

In this example, only one calculation step is required.



As you can see, the solution to this problem involves a balanced chemical equation and a mole to mole calculation. Step 2 and Step 4 are not required.

$$2\ Al_{(s)} + 3\ CuCl_{2(aq)} \rightarrow 2\ AlCl_{3(aq)} + 3\ Cu_{(s)}$$

$n = ?$ 1.65 mol

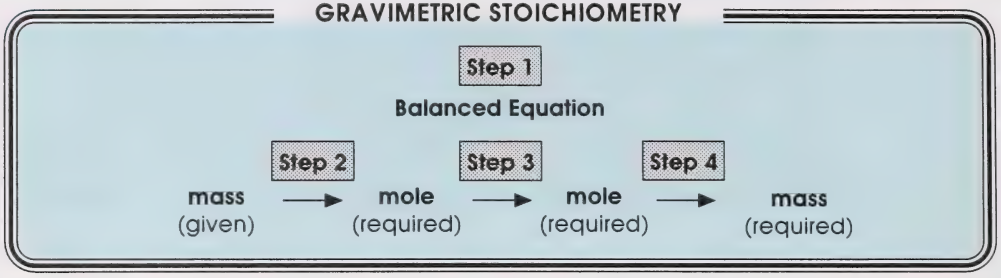
$$n_{(Al)} = 1.65\ \cancel{mol\ Cu} \times \frac{2\ mol\ Al}{3\ \cancel{mol\ Cu}}$$

$$= 1.10\ mol\ Al$$

Example 2

What mass of precipitate would be formed from the complete reaction of 0.68 mol of aqueous aluminum sulphate with aqueous calcium nitrate?

9. In the following graphic, circle the steps required to solve Example 2.



You can write the solution for Example 2 as follows.

$$Al_2(SO_4)_{3(aq)} + 3\ Ca(NO_3)_{2(aq)} \rightarrow 3\ CaSO_{4(s)} + 2\ Al(NO_3)_{3(aq)}$$

0.65 mol $\qquad\qquad$ $m = ?$

Since the calcium sulphate is the only solid product, it must be the precipitate.

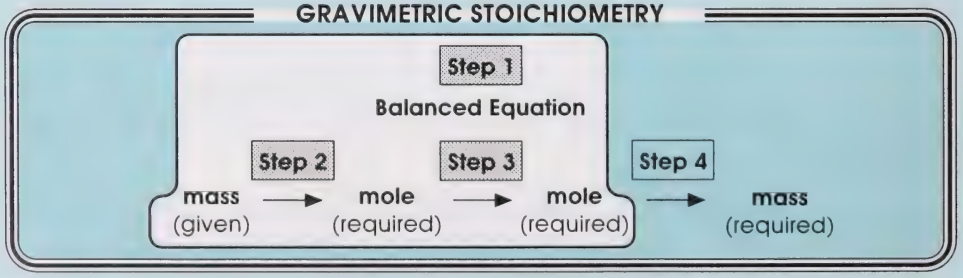
$$m_{(CaSO_4)} = 0.65\ \cancel{mol\ Al_2(SO_4)_3} \times \frac{3\ \cancel{mol\ CaSO_4}}{1\ \cancel{mol\ Al_2(SO_4)_3}} \times \frac{136.14\ g}{1\ \cancel{mol\ CaSO_4}}$$

$= 265\ g$ ← This answer must be rounded to two significant digits because of the 0.65 mol used.

$= 2.7 \times 10^2\ g$

Example 3

Another possible stoichiometry situation could involve the steps as highlighted in the following graphic.



10. Write an example question that would fit with the steps highlighted by the preceding graphic for Example 3. Use the following reaction as the basis for your question: Magnesium metal will react vigorously with oxygen to produce a white, ionic solid.

11. What two steps were common to all the previous stoichiometry examples?

12. When calculating a problem that involves the molar mass of one of the chemicals, is the coefficient of that chemical from the equation included in the molar mass? (Hint: Refer to Example 2 for help in answering the question.)

Check your answers by turning to the Appendix, Section 1: Activity 2.

The best way for you to gain confidence in solving stoichiometry problems is by practising. Answer the following questions which give you a variety of stoichiometry situations.

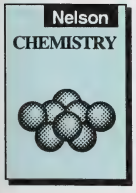
13. How many moles of oxygen are used when 0.951 mol of methane undergoes complete combustion?

14. What mass of nitrogen reacting in the cylinder of a car engine will react with oxygen to produce 15.07 mol of the pollutant nitrogen monoxide?

15. If 7.36 g lithium metal reacts with nitrogen gas, how many moles of $Li_3N_{(s)}$ are produced?

16. In a double replacement reaction called the Thermite reaction, aluminum reacts in a spectacular way with iron(III) oxide. The heat produced from this reaction is enough to melt the iron which is then used to weld steel together. What mass of solid iron(III) oxide is required for every 25.0 g of aluminum used?

Check your answers by turning to the Appendix, Section 1: Activity 2.



The stoichiometric method you have been using has allowed you to calculate on paper the required amounts of reactants or products from given measurements. But how would you test to see if your calculations apply to an actual chemical reaction in a lab? Read the bottom paragraph on page 171 and the top paragraph on page 172 of your textbook for an introduction to testing the stoichiometric method.

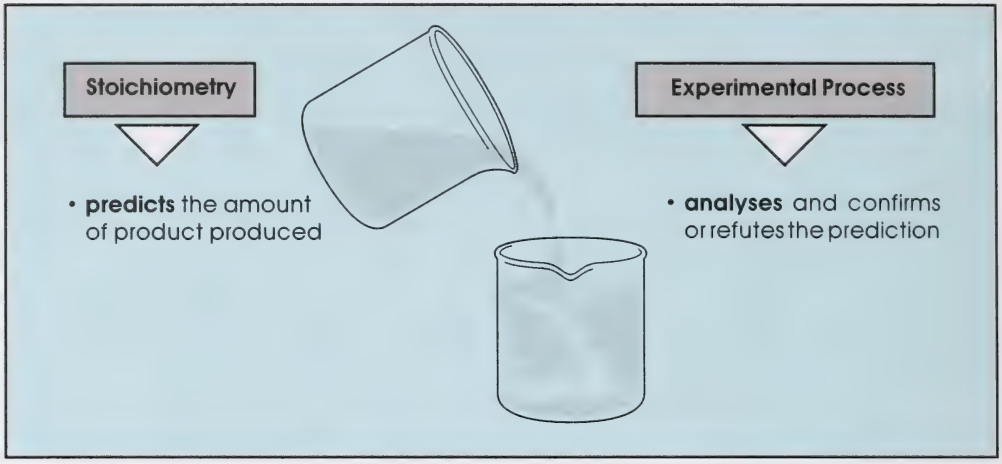
17. What is the most rigorous test of any scientific concept?

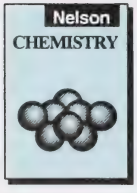
18. Under which conditions do you consider the prediction to be falsified?

19. a. What is a common experimental design for testing stoichiometric predictions?

    b. What is the relationship between the stoichiometry and the experimental design and process?

Check your answers by turning to the Appendix, Section 1: Activity 2.


Stoichiometry
• predicts the amount of product produced
Experimental Process
• analyses and confirms or refutes the prediction



As you discovered in Module 1, your textbook provides you with a good explanation on pages 522-527 of the scientific processes involved in experiments.

At this point, you are interested in the evaluation process in order to test the stoichiometric method. Review point number two on page 524.

The **percent difference** (or percent error) allows you to calculate how close your experimental (actual) results are to your predicted (or theoretical) results.

percent difference – the difference between the predicted value and the experimental or actual value, expressed as a percentage of the predicted value

$$\% \text{ difference} = \frac{|\text{experimental} - \text{predicted}|}{\text{predicted}} \times 100$$

20. In the preceding percent difference formula, what do the vertical lines in the formula mean?

. . . |experimental – predicted| . . .

21. Why is the $\times 100$ in the formula?

22. What value for the percent difference would tell you that your prediction was valid?

Check your answers by turning to the Appendix, Section 1: Activity 2.

Lab Exercise 7B: Testing Gravimetric Stoichiometry

Nelson CHEMISTRY

The preceding information on evaluation will allow you to complete Lab Exercise 7B on page 172 of your textbook. The following guidelines and questions will help you complete the lab exercise.

Guidelines

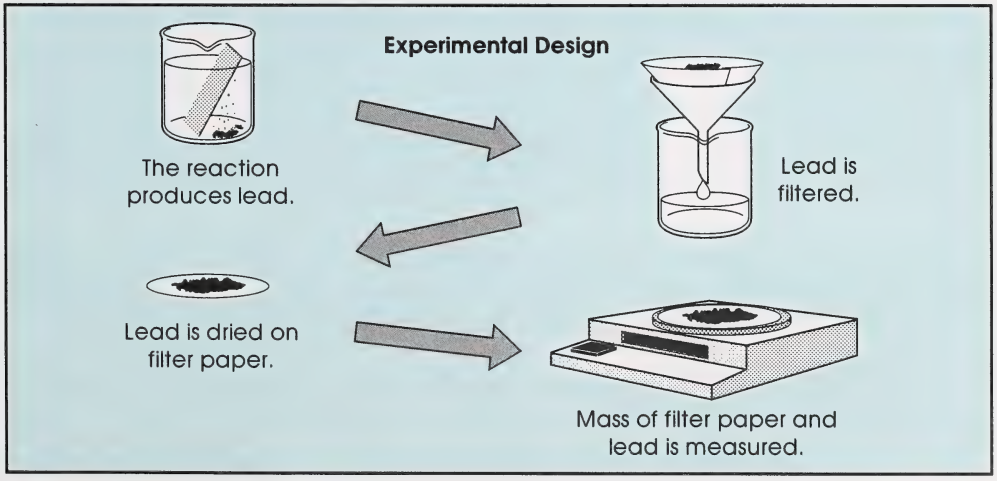
Science Skills

- [x] A. Initiating
- [ ] B. Collecting
- [ ] C. Organizing
- [x] D. Analysing
- [ ] E. Synthesizing
- [x] F. Evaluating

- The prediction required is a stoichiometric calculation, using the information supplied in the problem.
- Use the information on pages 522-527 of your textbook to help you with the other requirements of the lab exercise.
- The following questions will lead you through the lab exercise requirements.

23. Write a balanced chemical equation based on the information given in the problem. Write the given or known information from the problem under the equation.

24. Complete the prediction by doing the stoichiometry necessary to find the predicted mass of lead.


Experimental Design
The reaction produces lead.
Lead is filtered.
Lead is dried on filter paper.
Mass of filter paper and lead is measured.

25. Why is it important to know the mass of the filter paper?

26. What was the mass of lead produced?

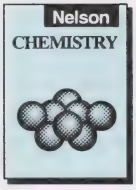
27. a. Write an evaluation of the experiment as instructed in the lab exercise. Include a percent difference calculation.

    b. Based on your answer to a., was your prediction valid? Explain.

Check your answers by turning to the Appendix, Section 1: Activity 2.

You will find that one of the most gratifying things to do in science is to design your own investigation and then perform that experiment to a successful completion. You will also find that becoming good at using the scientific process takes practice. In Activity 3, you will design and perform an investigation to test the stoichiometric method.

Lab Exercise 7C: Designing and Testing Stoichiometry



In order to give you more practice in experimental design and other processes before you do your own investigation, do Lab Exercise 7C on page 172 of your textbook. This lab exercise will give you more confidence and ability in scientific skills. The following questions will lead you through the lab exercise. If you need a review of any components of an investigation, refer to page 19 or pages 522-527 in your textbook.

Science Skills

- [x] A. Initiating
- [ ] B. Collecting
- [x] C. Organizing
- [x] D. Analysing
- [ ] E. Synthesizing
- [x] F. Evaluating

28. What is the name of the procedure required for making the prediction in Lab Exercise 7C?

29. a. Remember the use of the solubility table as instructed in Module 1. Which product is the precipitate in this reaction?

    b. Predict how much of the precipitate will form from 2.98 g of sodium phosphate.

30. Give a brief description of the experimental design for this lab exercise.

31. List the materials required for this experiment.

32. List the steps in your procedure.

33. Write a brief analysis of the investigation.

34. Write an evaluation for this lab exercise. Be sure you address each component of the evaluation.

Check your answers by turning to the Appendix, Section 1: Activity 2.

In this activity you have seen that the use of proportions to calculate quantities using chemical reactions is called stoichiometry. You have had more practice with problem solving and you should have discovered the value of stoichiometry in the scientific process. Your skills in problem solving will help you throughout the rest of this course.

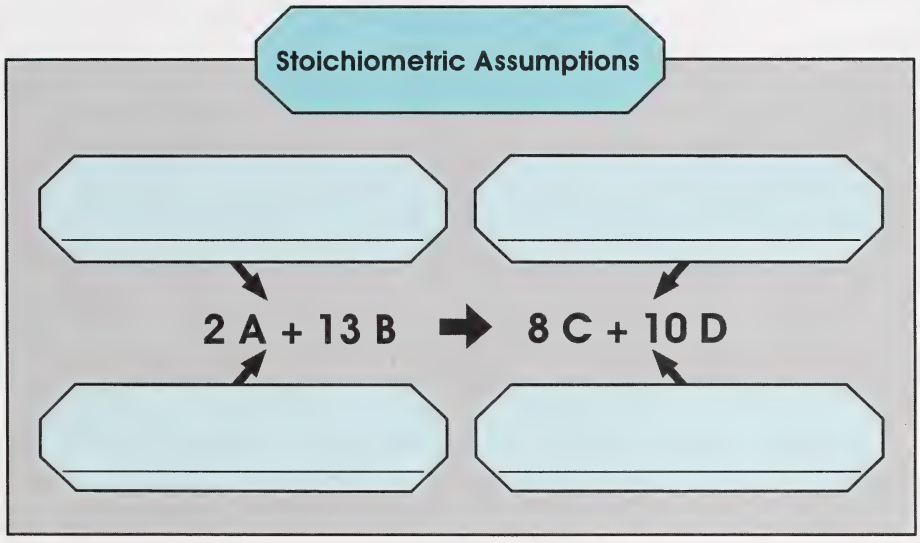
ACTIVITY 3

Testing Gravimetric Stoichiometry

assumption – accepting an idea to be true without proof

What is an assumption? Do assumptions make your life easier or harder? Believe it or not, in solving stoichiometry problems, several assumptions were made.

1. Refer to the top half of page 174 in your textbook and complete the following figure which refers to assumptions in stoichiometry.



Check your answers by turning to the Appendix, Section 1: Activity 3.

It is important to realize what assumptions are being made in any scientific process. This knowledge allows you to know when to be skeptical about a process, or to know when more investigation is necessary.

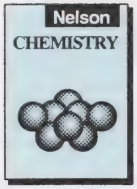
The assumptions you make in stoichiometry should allow you to perform the next investigation without any surprises.

Investigation 7.3: Testing the Stoichiometric Method

PATHWAYS

If you have access to laboratory facilities, do Part A.
If you do not have access to laboratory facilities, do Part B.

Part A



Carefully follow the directions given for Investigation 7.3 on page 174 of *Nelson Chemistry*. Pay special attention to the required components, safety aspects, and applied science skills.

Science Skills

- [x] A. Initiating
- [x] B. Collecting
- [x] C. Organizing
- [x] D. Analysing
- [ ] E. Synthesizing
- [x] F. Evaluating

Guidelines

- The copper(II) sulphate described in the investigation is usually supplied as copper(II) sulphate pentahydrate. The mass cited in the problem, 2.56 g, is enough to make excess of either copper(II) sulphate or copper(II) sulphate pentahydrate.
- Remember to measure the mass of the filter paper before using the paper.

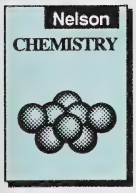
The following questions will lead you through the investigation.

2. Predict how much of the precipitate will form from 2.00 g of strontium nitrate.
3. Give a brief description of the experimental design for this investigation.
4. List the materials you will require for this investigation.
5. Write a procedure for this investigation. Remember to list the steps.
6. Record your evidence in an organized manner.
7. Write a brief analysis of your results.
8. Write an evaluation of your investigation, including percent difference.

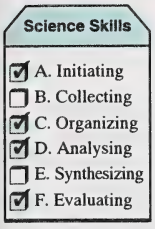
Check your answers by turning to the Appendix, Section 1: Activity 3.

End of Part A

Part B



Carefully read the directions given for Investigation 7.3 on page 174 of *Nelson Chemistry*. Pay special attention to the required components, safety aspects, and applied science skills. Review the guidelines given in Part A.



The following questions will lead you through the investigation.

9. Predict how much of the precipitate will form from 2.00 g of strontium nitrate.

10. Give a brief description of the experimental design for this investigation.

11. List the materials you will require for this investigation.

12. Write a procedure for this investigation. Remember to list the steps.

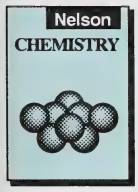
13. In order that you can do the analysis, check the Appendix for the answer to question 6, Section 1, Activity 3. Copy the given evidence chart onto your paper.

14. Write a brief analysis of the results.

15. Write an evaluation of the investigation, including percent difference.

Check your answers by turning to the Appendix, Section 1: Activity 3.

End of Part B



You can use stoichiometric calculations as analytical tools in many chemical processes. Read the section Stoichiometry in Chemical Analysis on page 174 of your textbook as an introduction to this concept.

In the following question, gravimetric stoichiometry is used as a strategy to find unknown information about substances.

16. Complete questions 15, 16, and 17 from pages 174-175 of your textbook.

Check your answers by turning to the Appendix, Section 1: Activity 3.

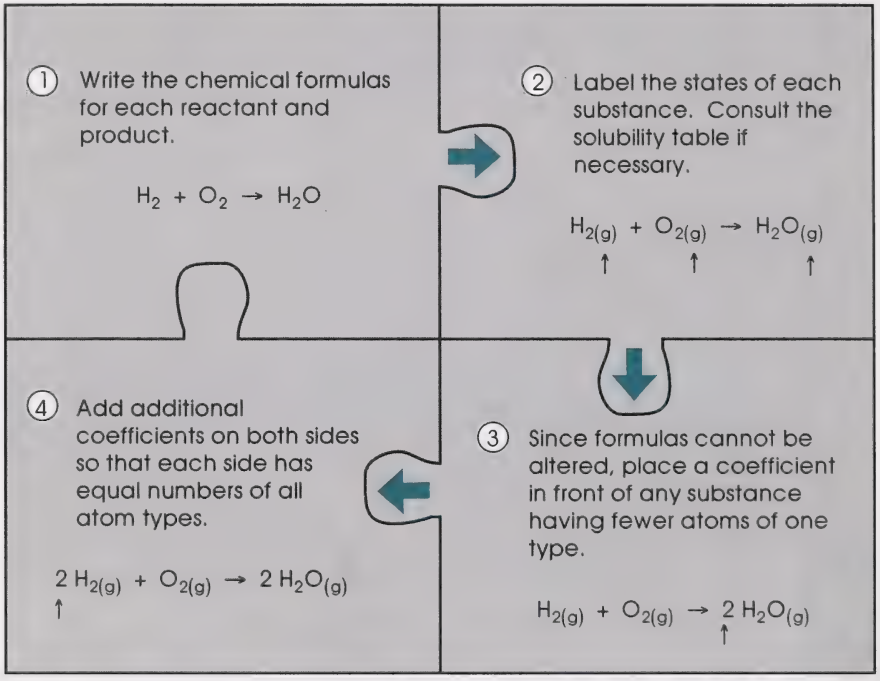
It is often better if you can test a process for yourself. In this activity, you have had a chance to test the stoichiometric method and demonstrate that its assumptions are valid.

Follow-up Activities

If you had difficulties understanding the concepts in the activities, it is recommended that you do the Extra Help. If you have a clear understanding of the concepts, it is recommended that you do the Enrichment.

Extra Help

Balancing Chemical Equations


1 Write the chemical formulas for each reactant and product.
$H_2 + O_2 \rightarrow H_2O$
2 Label the states of each substance. Consult the solubility table if necessary.
$H_{2(g)} + O_{2(g)} \rightarrow H_2O_{(g)}$
3 Since formulas cannot be altered, place a coefficient in front of any substance having fewer atoms of one type.
$H_{2(g)} + O_{2(g)} \rightarrow 2\ H_2O_{(g)}$
4 Add additional coefficients on both sides so that each side has equal numbers of all atom types.
$2\ H_{2(g)} + O_{2(g)} \rightarrow 2\ H_2O_{(g)}$

1. Write balanced chemical equations for each of the following.

   a. ___ $H_2S_{(g)} \rightarrow$ ___ $H_{2(g)}$ + ___ $S_{8(g)}$

   b. potassium iodide solution + lead(II) nitrate solution →

   c. fluorine gas + sodium hydroxide solution → sodium fluoride solution + oxygen gas + water

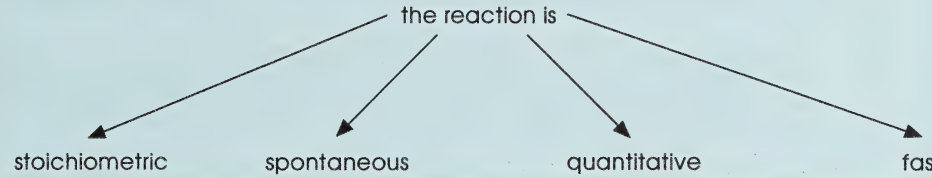
   d. A precipitate is formed when solutions of iron(III) nitrate and ammonium hydroxide are mixed.

Gravimetric Stoichiometry

Gravimetric stoichiometry deals with the following:

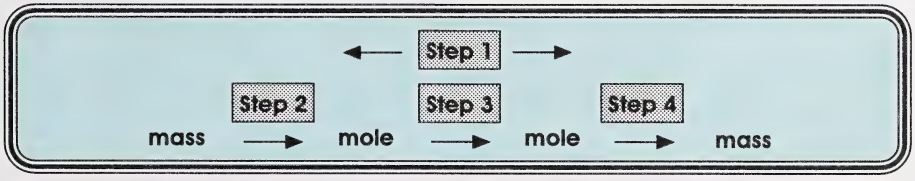
- relates the masses and moles of reactants and products in a chemical reaction
- uses molar mass values and mole ratios as conversion factors

Assumptions made in using this stoichiometric method are that



2. You are given the following problem: What mass of solid silver can be plated out when 16.2 g of copper metal is completely reacted in a solution of silver nitrate?

   a. Write the appropriate step numbers in the blanks of the given statements.



- _____ convert moles of given to moles required
- _____ write the balanced chemical equation
- _____ convert moles required to mass required
- _____ convert mass given to moles given

   b. Solve the given problem.

3. Which steps from the figure given in question 2. a. would be needed to solve the following problems?

   a. In the Haber process, how many grams of ammonia can be formed from the reaction of 26.8 mol of hydrogen gas with excess nitrogen gas?

   b. If 468 g of propane, $C_3H_{8(g)}$, burned completely in your barbecue, how many moles of carbon dioxide gas would form?

4. Solve the problems given in 3. a. and 3. b.

Remember the following:

- The only time coefficients from the balanced equation are used in a stoichiometry problem is in the mole ratio. The coefficients are never factored into the molar mass.
- Always keep track of units – they serve as checks for showing you whether or not you have set up the calculations properly.

Check your answers by turning to the Appendix, Section 1: Extra Help.

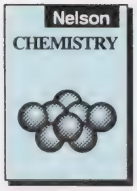
Enrichment

Do either **or** both of the following questions.

1. You may have seen a volcano in a movie. Special effects technicians are able to make miniature volcanoes by decomposing ammonium dichromate. The orange crystalline substance is decomposed to produce green chromium(III) dioxide, water vapour, and nitrogen gas. If the special effects technician started by heating 3 g of ammonium dichromate, what is the mass of each of the products when the reaction is complete?

2. Much of what you have examined to this point has involved stoichiometry on a small scale. Some chemical technicians use large-scale stoichiometry every working day.

   Complete question 30 **or** 32 from pages 198-199 of your textbook.



Check your answers by turning to the Appendix, Section 1: Enrichment.

Conclusion

In this section you reviewed the concept of the mole and how chemists use the concept to deal with a large number of particles. You also reviewed the balancing of chemical equations. These concepts are fundamental to stoichiometry, which describes the quantitative relationships that can be solved through a chemical equation. You also used stoichiometric calculations to predict the quantities of products formed in several investigations.

50

Section 1 Assignment: Introduction to Stoichiometry

Review the Evaluation information found in the introductory pages of this module.

It is important to number and clearly identify each page with the following information at the top:

Chemistry 20 – Module 3 Section 1 Assignment Page # Name and ID #

Be sure to write legibly. Leave a wide left margin and number all of your pages.

1. Write balanced chemical equations, including states, for each of the following reactions. **(6 marks)**

   a. Solid potassium chlorate can be very dangerous if stored near flammable materials. Potassium chlorate decomposes upon heating to form solid potassium chloride and oxygen gas.

   b. Reacting a solution of aluminum sulphate with aqueous lead(II) nitrate will remove the toxic lead ions from solution.

2. For the following problem, find and list three errors in the given solution. Show the corrections. **(6 marks)**

   Problem

   How many grams of calcium carbonate would be needed to react completely with 15.0 g of hydrochloric acid to produce carbon dioxide, calcium chloride, and water?

   Solution

   $$\begin{array}{lll} CaCO_{3(s)} & + \quad 2\,HCl_{(aq)} \rightarrow & CO_{2(g)} + CaCl_{2(aq)} + 2\,H_2O_{(l)} \\ m = ? & 15.0\ g & \\ 100.09\ g/mol & 36.46\ g/mol & \end{array}$$

   $$m_{(CaCO_3)} = 15.0\ g\ HCl \times \frac{1\ mol\ HCl}{36.46\ g\ HCl} \times \frac{1\ mol\ CaCO_3}{2\ mol\ HCl} \times \frac{1\ mol\ CaCO_3}{100.09\ g}$$
   $$= 0.00206\ g\ CaCO_3$$

   The amount of calcium carbonate required is 0.00206 g.

3. Find the number of moles each of the following represents. Show your work. **(6 marks)**

   a. 6.28 g ammonium chloride

   b. 2.5 kg sodium silicate

   c. 3.16 mg lead(II) sulphide

4. a. A standard basket of fruit from a fruit stand contains two apples and three oranges. If you were delivering a van with enough fruit to make up 20 baskets, how many apples and oranges would you put in the van? **(2 marks)**

   b. Relate the proportions problem in 4. a. to the mole ratio in a balanced equation. **(2 marks)**

5. Write a word description of the following equation in terms of atoms and molecules, not moles. **(2 marks)**

$$C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$$

6. Solid calcium nitride will react with water to form ammonia gas and solid calcium hydroxide. How many grams of calcium nitride must completely react to form 4.60 mol of ammonia? Show your work. **(6 marks)**

7. How many grams of magnesium metal will completely react with 50.0 g of phosphoric acid (aqueous hydrogen phosphate)? Show your work. **(7 marks)**

Science Skills

- [x] A. Initiating
- [ ] B. Collecting
- [ ] C. Organizing
- [x] D. Analysing
- [ ] E. Synthesizing
- [x] F. Evaluating

8. A student wished to test the stoichiometric method using a precipitation-filtration experimental design.

   The problem of the investigation was stated as follows:

   What is the mass of precipitate formed when 5.00 g of copper metal reacts completely with a solution of silver nitrate?

   The following evidence was collected:

   - Shiny solid crystals formed on the surface of the copper, but fell to the bottom of the beaker when the solution was shaken.
   - The solution turned a pale blue colour.
   - The mass of Cu metal was 5.00 g.
     The mass of filter paper plus precipitate was 19.07 g.

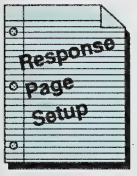
The percent difference was calculated to be 12.2% based on the 19.07 g experimental value of the filter paper plus precipitate and the predicted value calculated.

a. Complete the prediction of the investigation. Show all of your work. **(5 marks)**

b. What precipitate was formed? **(2 marks)**

c. Evaluate (in 1/2 to 3/4 page) the experimental results based on the following: **(6 marks)**

- Comment on the size of the percent difference.
- Comment on possible sources of error.
- Based on the results, would you change the experimental design? Explain.



Reproduce the science skills assessment box after your response to this question. Remember to indicate your evaluation of your skill level for each identified science skill. Be sure to include the spaces for teacher assessment.

Self:	A.	☐	B.	☐	C.	☐	D.	☐	E.	☐	F.	☐
Teacher:	A.	☐	B.	☐	C.	☐	D.	☐	E.	☐	F.	☐

SECTION 2

Applications of Stoichiometry

Have you ever added baking soda to vinegar and watched it fizz? Maybe you have even tried to blow up a balloon with the gas that is produced from that reaction. Could you measure the volume of gas produced or find the concentration of the vinegar used? Your knowledge of stoichiometry from the previous section will allow you to deal with similar situations in this section involving gases and solutions. The areas of gas and solution stoichiometry have wide ranges of applications.

What applications are there for stoichiometry in industry? Are the principles behind these applications different from the ones you have discovered already? In this section, you will also analyse several industrial processes which involve quantitative measurements of chemical reactions.

ACTIVITY 1

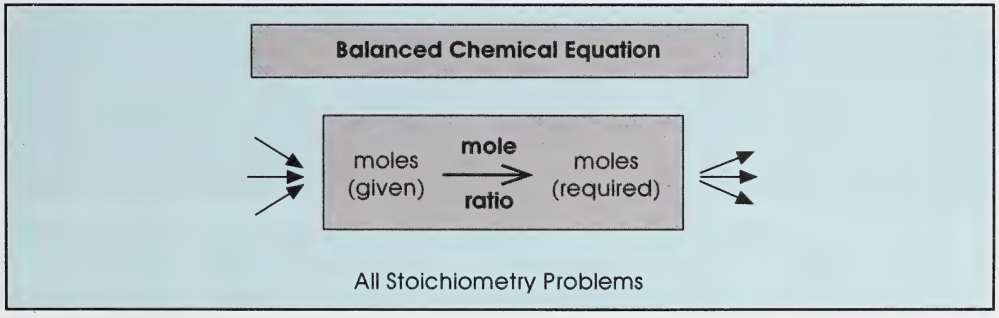
Working with Gases

You will be happy to learn that, even though the different terms gravimetric, gas, or solution stoichiometry may be used, there is a common method of problem-solving behind all of them. Only the types of measurements or units used change in the various methods.

1. What steps do you think will be common between all the types of stoichiometric methods?

Check your answers by turning to the Appendix, Section 2: Activity 1.

Having these common stoichiometric steps means that, no matter what measurements you are given, and no matter what answers you are required to find, you are always working with moles in your calculations. The mole ratio will always provide the bridge between given quantities and required answers.



In the previous section, when you solved gravimetric stoichiometry problems, you used the molar mass to convert the given mass into moles, as shown in Figure 3.1.

In this activity, gas stoichiometry may involve different given and required units. Gases are often measured by volume, not mass. Therefore, instead of converting mass to moles, you may need to convert the volume of a gas to moles.

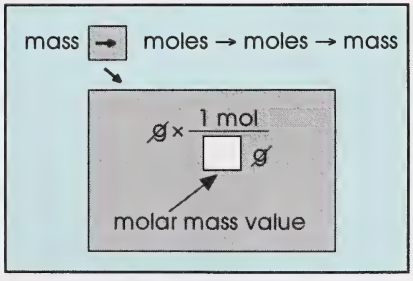


FIGURE 3.1: Mass to Mole Conversion

2. What two units will have to be in the conversion factor that changes volume to moles?

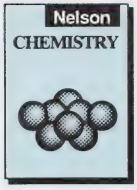
$$L \times \boxed{?} = mol$$

Check your answers by turning to the Appendix, Section 2: Activity 1.

In Module 2, you worked with gases and the variables that describe gases – pressure, volume, temperature, and moles. These variables will now be considered along with chemical reactions involving gases.

molar volume – the number of litres in volume occupied by one mole of a gas

For example, in a problem involving the conversion of gas volume to moles, as discussed previously, you would need to know the **molar volume** of a gas. The molar volume depends on the conditions of temperature and pressure.



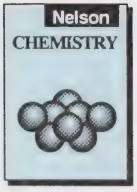
Discover the way that gases are dealt with in chemical reactions by reading from the bottom of page 175 to the end of page 177 of the textbook. Study the examples carefully.

STP Conditions	SATP Conditions	Other Temperature and Pressure
↓	↓	Ideal Gas Law ↓
molar volume v = 2.44 L/mol	molar volume v = 24.8 L/mol	molar volume v = ?

3. To what kinds of chemical reactions does the stoichiometric method apply?

4. Compare the general steps of the stoichiometric calculations for solids and gases.

Check your answers by turning to the Appendix, Section 2: Activity 1.



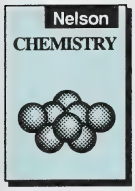
Turn to page 176 of your textbook. Answer the following two questions which pertain to the given problem on that page.

5. Why do you need to convert the 300-g mass of propane to moles?

Remember you can use a conversion factor or its reciprocal, whichever allows you to cancel the units to get the correct answer.

6. Why do you use L/mol as the units in the last conversion factor?

Check your answers by turning to the Appendix, Section 2: Activity 1.



You discovered in the previous section that unit analysis allows you to do all the calculations on one line as long as you keep track of the units. You could do the problem illustrated on the top of page 177 of *Nelson Chemistry* using unit analysis.

$$2\,Na_{(s)} + 2\,HOH_{(l)} \rightarrow H_{2(g)} + 2\,NaOH_{(aq)}$$

$$m_{(Na)} = 20.0\ \cancel{L\,H_2} \times \frac{1\ \cancel{mol\ H_2}}{24.8\ \cancel{L\,H_2}} \times \frac{2\ \cancel{mol\ Na}}{1\ \cancel{mol\ H_2}} \times \frac{22.99\ g}{1\ \cancel{mol\ Na}}$$

$$= 37.1\ g$$

You may use the method for calculations that you are most comfortable with. In the rest of this section, the examples will be shown using the style shown in the textbook.

FIGURE 3.2: Gas Production from Zinc and Antacid Tablets in Hydrochloric Acid

Why do you often have to burp when you take an antacid tablet? A possible clue can be seen in the right-hand beaker in Figure 3.2. Is it possible to calculate the volume of gas produced in a chemical reaction? Consider the following example.

What is the volume of $H_{2(g)}$ produced at STP when 27.53 g zinc is completely reacted in a hydrochloric acid solution, as illustrated in Figure 3.2?

$$Zn_s + 2\,HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$$

27.53 g — v = ?

65.38 g/mol — 22.4 L/mol

$$n_{(Zn)} = 27.53\ \cancel{g\ Zn} \times \frac{1\ mol\ Zn}{65.38\ \cancel{g\ Zn}}$$

$$= 0.4211\ mol\ Zn$$

$$n_{(H_2)} = 0.4211 \text{ mol Zn} \times \frac{1 \text{ mol } H_2}{1 \text{ mol Zn}}$$

$$= 0.4211 \text{ mol } H_2$$

$$v_{(H_2)} = 0.4211 \text{ mol } H_2 \times \frac{22.4 \text{ L}}{1 \text{ mol } H_2}$$

$$= 9.43 \text{ L}$$

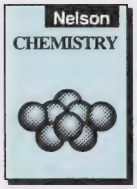
or,

in one line

$$v_{(H_2)} = 27.38 \text{ g Zn} \times \frac{1 \text{ mol Zn}}{65.38 \text{ g Zn}} \times \frac{1 \text{ mol } H_2}{1 \text{ mol Zn}} \times \frac{22.4 \text{ L}}{1 \text{ mol } H_2}$$

$$= 9.43 \text{ L}$$

To gain experience in performing gas stoichiometry calculations, do the following questions.



7. Complete questions 18 and 19 (methanol is CH_3OH) from page 178 of your textbook.

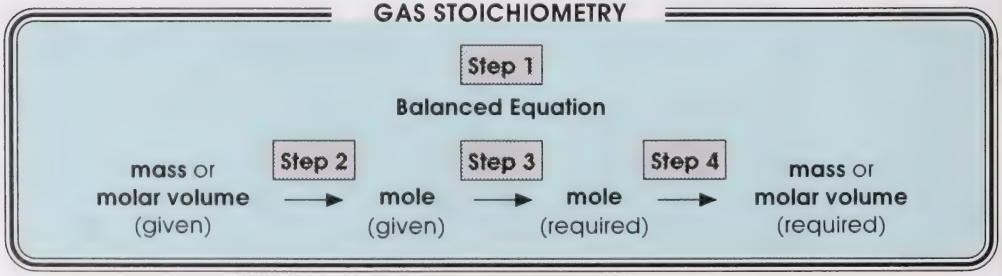
8. Sulphur dioxide gas is a pollutant which is sometimes recovered to make sulphuric acid. The first step in this process can be described by the balanced equation:

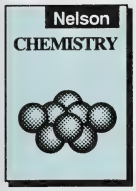
$$2\,SO_{2(g)} + O_{2(g)} \rightarrow 2\,SO_{3(g)}$$

How many litres of oxygen would be required to react with 15.75 L of sulphur dioxide at STP?

Check your answers by turning to the Appendix, Section 2: Activity 1.

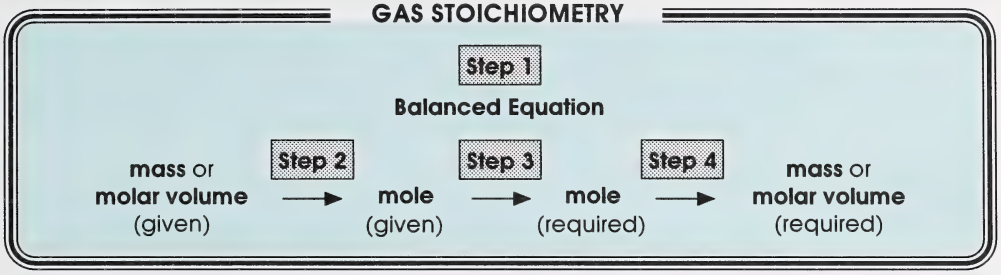
Stoichiometric methods involving gases can be summarized as follows:





In some cases, problems are not stated in standard conditions of STP or SATP. In Module 2 you saw how to solve these situations using the Ideal Gas Law. Now you will examine how to use the Ideal Gas Law along with chemical reactions in a stoichiometric situation. Work through the example on the bottom of page 177 in your textbook to see how this type of problem is solved.

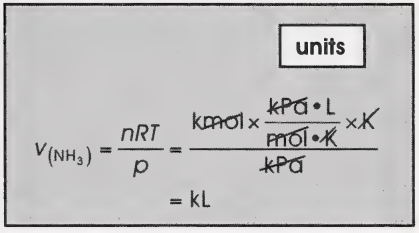
9. a. In the following figure, circle the step of the calculation that is replaced by the Ideal Gas Law calculation in the textbook example.



b. Where else may the Ideal Gas Law be used in a gas stoichiometry calculation?

Check your answers by turning to the Appendix, Section 2: Activity 1.

Notice in the box on the right that even though you are using a formula to solve part of the problem, the units are still properly arranged to give the correct unit for the answer.



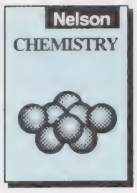
Practise using the Ideal Gas Law in stoichiometric calculations by doing the following question.

10. Some cars are equipped with air bags that work by decomposing sodium azide (NaN_3) and forming nitrogen gas when an accident occurs, as shown in the following equation:

$$2\,NaN_{3(s)} \rightarrow 2\,Na_{(s)} + 3\,N_{2(g)}$$

Calculate the volume of $N_{2(g)}$ formed at 20°C and 104.8 kPa from 58.0 g of NaN_3.

Check your answers by turning to the Appendix, Section 2: Activity 1.



You have already done several problems in Module 2 and in this activity dealing with the universal gas constant. What kind of variables are necessary to calculate the universal gas constant (R) in stoichiometric problems? To answer this question, do Investigation 7.4 on pages 178-179 of your textbook.

Investigation 7.4: The Universal Gas Constant

PATHWAYS

If you have access to laboratory facilities, do Part A.
If you do not have access to laboratory facilities, do Part B.

Science Skills

- ☐ A. Initiating
- ☑ B. Collecting
- ☑ C. Organizing
- ☑ D. Analysing
- ☐ E. Synthesizing
- ☑ F. Evaluating

Part A

Carefully follow the directions given for this investigation. Pay special attention to the required components, safety aspects, and applied science skills.

Guidelines

- You must wear safety glasses and a lab apron throughout the entire experiment.
- Depending on the thickness of the magnesium ribbon, the amount suggested in *Nelson Chemistry* may be too great. If the measured mass is greater than 0.05 g, you should use a strip about 40-50 mm in length.

Evaluation

11. Record your evidence in a table.

Analysis

12. For your analysis, calculate the universal gas constant R, using the data you collected.

13. Evaluate the experimental design, including the percent difference between your experimental value and the predicted (accepted) value of 8.31 kPa•L/mol•K for R. Discuss two possible sources of error in this experiment.

Check your answers by turning to the Appendix, Section 2: Activity 1.

End of Part A

Part B

Nelson CHEMISTRY

Carefully read the directions given for Investigation 7.4 on pages 178-179 of your textbook. Pay special attention to the required components, safety aspects, and applies science skills. Review the quidelines given in Part A.

Analysis

Science Skills

- [ ] A. Initiating
- [ ] B. Collecting
- [ ] C. Organizing
- [x] D. Analysing
- [ ] E. Synthesizing
- [x] F. Evaluating

14. In order that you can do the analysis, check the Appendix for the answer to question 11, Section 2, Activity 1. Copy the given evidence table onto your paper.

15. For your analysis, calculate the universal gas constant R, using the data you were given.

16. Evaluate the experimental design, including the percent difference between the experimental value and the predicted (accepted) value of 8.31 kPa•L/mol•K for R. Discuss two possible sources of error in this experiment.

Check your answers by turning to the Appendix, Section 2: Activity 1.

End of Part B

You have now seen that the stoichiometric method is a universal problem-solving tool. You have discovered the method's use in gravimetric and gas situations. Will the method also apply to solutions?

ACTIVITY 2

Quantities in Solution

1. From your work in Module 1, what variables can be involved when you describe substances in solution?

Check your answers by turning to the Appendix, Section 2: Activity 2.

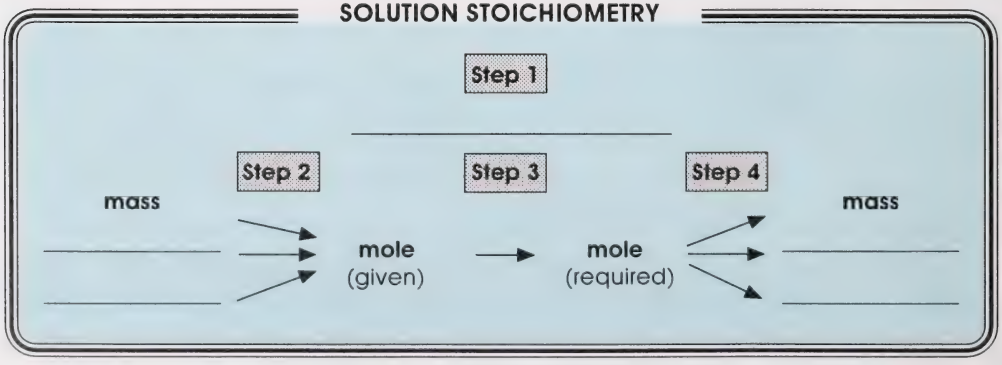
How will the solution variables be involved with reactions in solution? Will the stoichiometric method have to be altered to accommodate solution concepts? You will explore these questions in this activity.

Much of the work that is done in industry and scientific research makes use of stoichiometry and chemical reactions in solution. For further insight into solution stoichiometry, read pages 180 and 181 in your textbook. Pay special attention to the examples and the summary.

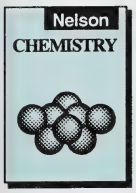
2. In solution stoichiometry, what values usually replace the mass and molar mass of a substance?

3. List the general steps required in all stoichiometry problems.

4. Fill in the blanks to represent the appropriate solution variables in the following figure.



Check your answers by turning to the Appendix, Section 2: Activity 2.



The example given on page 180 of your textbook describes part of a process mentioned previously for making fertilizer – the Haber process. This process is often called the Haber-Bosch process in honour of both contributing scientists.

This fertilizer example can be written in the one-line method as follows:

$$2\,NH_{3(aq)} + H_3PO_{4(aq)} \rightarrow (NH_4)_2HPO_{4(aq)}$$

$v = ?$ 12.9 mol/L

14.8 mol/L 1.00 ML

↑

stands for mega $(10^6\,L)$

$$v_{(NH_3)} = 1.00 \times 10^6\ \cancel{L} \times \frac{12.9\ \cancel{mol\ H_3PO_4}}{\cancel{L}} \times \frac{2\ \cancel{mol\ NH_3}}{1\ \cancel{mol\ H_3PO_4}} \times \frac{1\ L}{14.8\ \cancel{mol\ NH_3}}$$

$$= 1.74 \times 10^6\ L$$

$$= 1.74\ ML$$

The stoichiometry problems you have been calculating can also be solved using the formulas you used in Module 1.

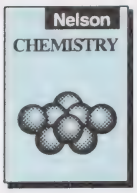
Step 1 Write a balanced chemical equation, the same as other methods.

Step 2 Convert to moles. Use $n = \frac{m}{M}$ or $n = Cv$.

Step 3 Find the required moles using mole ratio, the same as other methods.

Step 4 Find the required answer. Use $m = nM$ or $C = \frac{n}{v}$, $v = \frac{n}{C}$.

Practice with solution stoichiometry problems will give you more confidence in the process. Choose whatever calculation process best suits you to answer the following question.



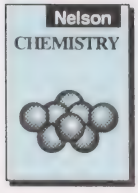
5. Complete questions 23, 24, and 25 from page 181 of your textbook. (Slaked lime is calcium hydroxide.)

Check your answers by turning to the Appendix, Section 2: Activity 2.

AOSTRA

Earlier in this module, you tested the ability of stoichiometry to make predictions in precipitation reactions. In the following lab exercise, you will be testing the ability of stoichiometry to make predictions relating to solutions.

Lab Exercise 7E: Testing Solution Stoichiometry



Do Lab Exercise 7E on page 182 of your textbook. Then complete the following questions.

6. Using solution stoichiometry, develop a prediction to answer the problem for Lab Exercise 7E.

7. Complete the experiment analysis. Calculate the mass of precipitate produced.

Science Skills

- [x] A. Initiating
- [ ] B. Collecting
- [ ] C. Organizing
- [x] D. Analysing
- [ ] E. Synthesizing
- [x] F. Evaluating

8. Evaluate the experimental design. Does this design test solution stoichiometry in such a way as to validly test the prediction you have made? Include a percent difference calculation.

Check your answers by turning to the Appendix, Section 2: Activity 2.

AOSTRA

In industry, chemical technicians rely on stoichiometric principles for making predictions. In the following exercise you will be using your knowledge of solution stoichiometry to determine the molar concentration of silver nitrate in a solution.

Lab Exercise 7F: Determining a Solution Concentration

Nelson CHEMISTRY

Do Lab Exercise 7F on page 182 of your textbook.

9. What is the lab test for silver nitrate that is implied in this exercise?

10. Write a brief analysis of the investigation.

Science Skills
- [ ] A. Initiating
- [ ] B. Collecting
- [ ] C. Organizing
- [x] D. Analysing
- [ ] E. Synthesizing
- [ ] F. Evaluating

11. Why do you think this lab exercise does not ask you to evaluate the experimental design?

Check your answers by turning to the Appendix, Section 2: Activity 2.

Nelson CHEMISTRY

So far you have been looking at many aspects of stoichiometry. These are all related, and at times you must combine the knowledge to make practical applications. Do Lab Exercise 7G on page 183 of your textbook.

Lab Exercise 7G: Solution and Gas Stoichiometry

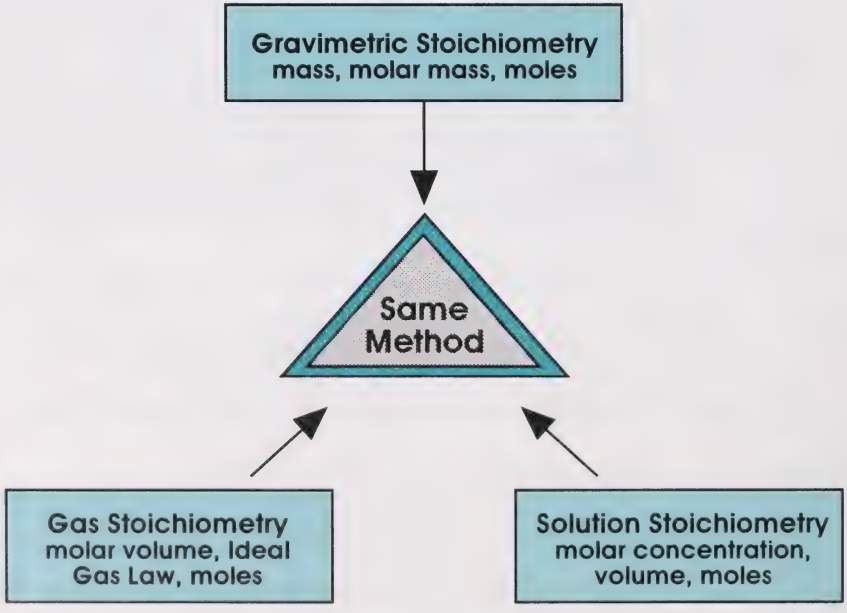
In Lab Exercise 7G, you will use both solution and gas stoichiometry in an analysis of baking.

Science Skills
- [x] A. Initiating
- [ ] B. Collecting
- [ ] C. Organizing
- [ ] D. Analysing
- [ ] E. Synthesizing
- [ ] F. Evaluating

12. Complete the prediction for this lab exercise.

Check your answers by turning to the Appendix, Section 2: Activity 2.

An important point for you to remember is that, even though different variables have been associated with gravimetric, gas, and solution stoichiometry, the method of stoichiometric problem-solving has generally been the same in each case.



ACTIVITY

3 Stoichiometry and Industry

Nelson CHEMISTRY

Can you imagine becoming famous for finding a way to make chalk and salt react? Ernest Solvay, a Belgian chemist, did precisely that in the 1860s. The Solvay process was consequently named after him. Read the case study of the Solvay process on pages 183-184 of your textbook.

1. Why is sodium carbonate such an important chemical on a worldwide basis?

2. What were the alternatives to the Solvay process in the last century?

3. How does the Solvay process separate the ammonium chloride from the sodium hydrogen carbonate?

Nelson CHEMISTRY

4. Complete question 33 (a, b, c) from page 185 of your textbook.

Check your answers by turning to the Appendix, Section 2: Activity 3.

There are many other examples of industrial applications of stoichiometry. You have already considered the Haber-Bosch process for making ammonia fertilizer.

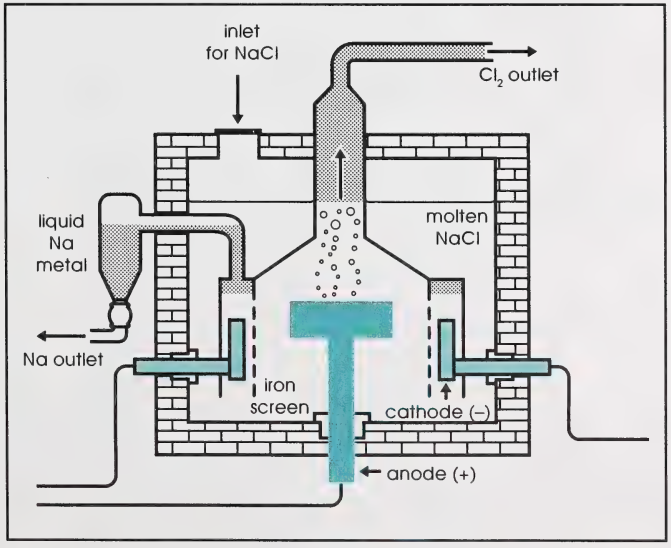
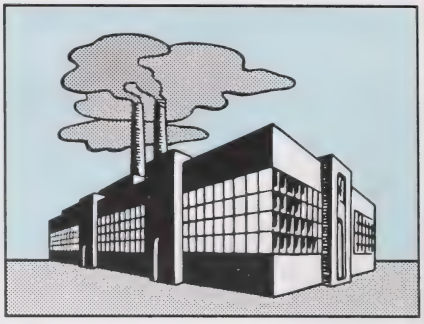


FIGURE 3.3: Down's Cell

Another important process for producing sodium metal and chlorine utilizes a simple decomposition reaction and high temperatures. This process occurs in a large chamber called a Down's Cell.

The chlorine produced from the Down's Cell can be used to produce hydrochloric acid for commercial use and bleach for home use. An industrialist using the Down's Cell would use stoichiometric methods to calculate the amounts of sodium chloride required to produce the desired quantities of sodium and chlorine products. The desired quantities of products would be based on the demand by society and industry.



Nelson CHEMISTRY

5. Complete question 21 from page 178 of your textbook.

Check your answers by turning to the Appendix, Section 2: Activity 3.

You have seen several examples of industrial applications of stoichiometry, paying special attention to the Solvay process and Down's Cell. You also compared empirical data with stoichiometric predictions in several investigations. In each case, the evidence gathered validated the stoichiometric method.

6. a. What if, in each investigation you studied, your percent difference had been greater than 10%?

   b. What considerations would you then make in assessing experimental designs and the stoichiometric method?

   c. Would just one investigation with a large percent difference be sufficient to question or change the stoichiometric method? Explain.

Check your answers by turning to the Appendix, Section 2: Activity 3.

Stoichiometry is a valuable tool that has many applications not only in chemistry, but in other areas as well. The problem-solving methods you have learned will help you in many aspects of your life.

Follow-up Activities

If you had difficulties understanding the concepts in the activities, it is recommended that you do the Extra Help. If you have a clear understanding of the concepts, it is recommended that you do the Enrichment.

Extra Help

General Stoichiometry Steps

Step 1 Write a balanced chemical equation. Write all given information under the appropriate substances in the equation.

↓

Step 2 Find the moles of the given substance. The given substance is the substance in the equation which has enough information given to calculate moles.

↓

Step 3 Use the mole ratio to convert moles given to moles required. The required substance is on top of the ratio and the numbers in the ratio are the coefficients from the equation.

↓

Step 4 Find the desired answer with its appropriate units.

The following should be noted:

- Steps 1 and 3 are common to all types of stoichiometric methods.
- Steps 2 and 4 may involve different variables specific to gravimetric, gas, and solution stoichiometry.

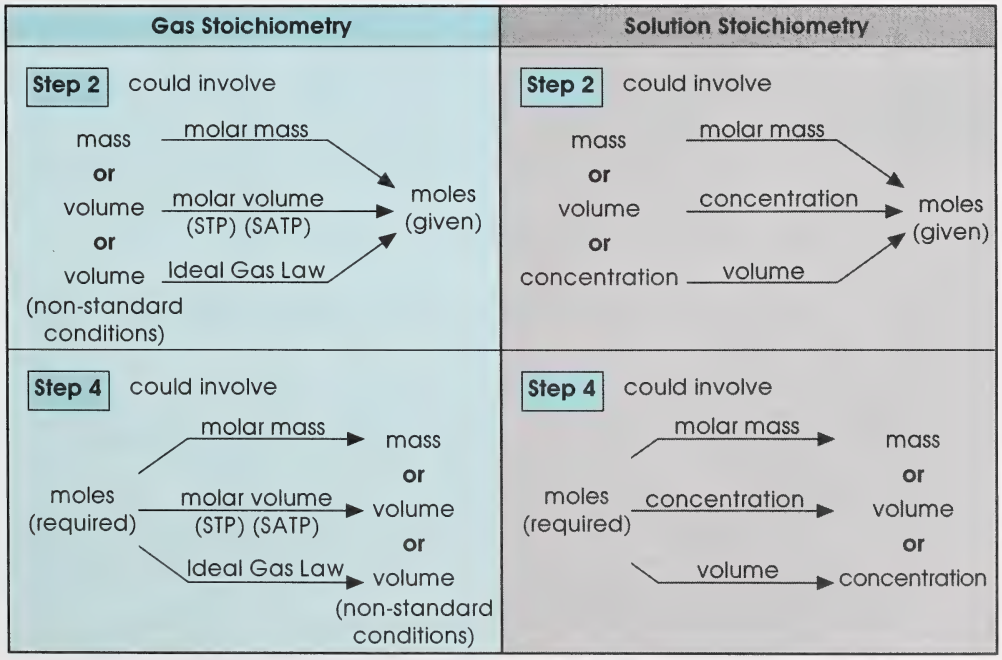


FIGURE 3.4: Steps in Gas and Solution Stoichiometry

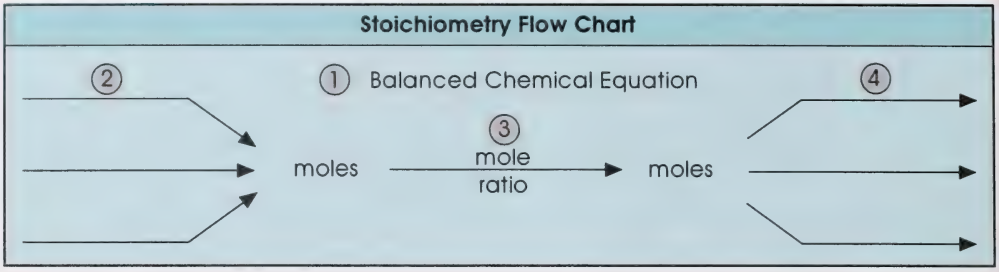
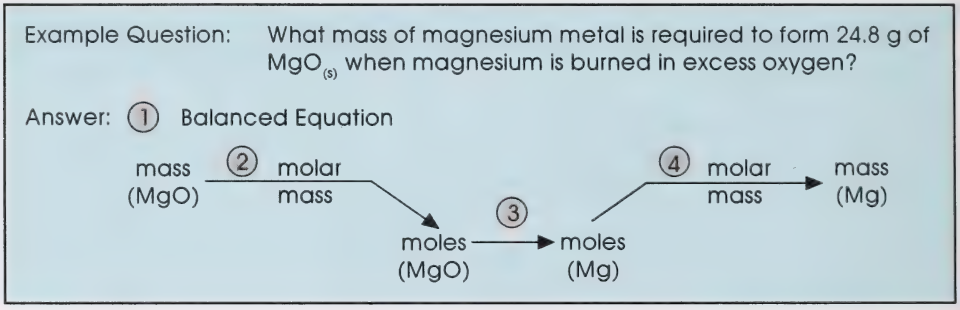


FIGURE 3.5: Flow Chart

1. For the following problems, use the flow chart from Figure 3.5 and the information from Figure 3.4. Draw a pathway for solving each of the following problems.


Example Question: What mass of magnesium metal is required to form 24.8 g of $MgO_{(s)}$ when magnesium is burned in excess oxygen?
Answer: 1 Balanced Equation
mass (MgO)
2 molar mass
moles (MgO)
3
moles (Mg)
4 molar mass
mass (Mg)

   a. Given the following equation, $C_2H_{4(g)} + H_{2(g)} \rightarrow C_2H_{6(g)}$, a 14.0 g sample of $C_2H_{4(g)}$ requires what volume (litres) of $H_{2(g)}$ at STP for a complete reaction?

   b. For the reaction, $2\,NH_{3(aq)} + H_2SO_{4(aq)} \rightarrow (NH_4)_2SO_{4(aq)}$, what volume of 6.00 mol/L $H_2SO_{4(aq)}$ is required to completely react with 50.0 mL of 2.50 mol/L $NH_{3(aq)}$?

2. Match the conversion factors from the left column that could be used to solve the problem statements in the right column.

a. 22.4 L/mol	i. Hydrogen gas is made by reacting steam with methane. The resulting hydrogen is measured at 25°C and 120 kPa.
b. 24.8 L/mol	ii. Water is decomposed into hydrogen gas and oxygen gas. The volume of the gases is measured at SATP.
c. Ideal Gas Law	iii. The toothpaste ingredient tin(II) fluoride is formed by reacting the solid metal with hydrogen fluoride at STP.

3. Look at the solutions to the following problems. One or two terms are incorrect in each solution. Write the correct term(s) in the appropriate boxes below each solution.

a. $H_2SO_{4(aq)} + 2\,NaOH_{(aq)} \rightarrow 2\,H_2O_{(l)} + Na_2SO_{4(aq)}$

0.100 mol/L	150 mL
v = ?	0.020 mol/L

Calculate the volume required.

$$v = 150\text{ mL NaOH} \times \frac{0.020\text{ mol NaOH}}{1\text{ L NaOH}} \times \frac{2\text{ mol NaOH}}{1\text{ mol } H_2SO_4} \times \frac{1\text{ L } H_2SO_4}{0.100\text{ mol } H_2SO_4}$$

[ ] × [ ] × [ ] × [ ]

b. $Pb(NO_3)_{2(aq)} + Na_2SO_{4(aq)} \rightarrow 2\,NaNO_{3(aq)} + PbSO_{4(s)}$

v = 377 mL	48.5 mL
C = ?	0.050 mol/L

Calculate the concentration of $Pb(NO_3)_{2(aq)}$.

$$C = 48.5\text{ mL } Na_2SO_4 \times \frac{1\text{ L}}{0.050\text{ mol } Na_2SO_4} \times \frac{1\text{ mol } Pb(NO_3)_2}{1\text{ mol } Na_2SO_4} \times 377\text{ mL } Pb(NO_3)_2$$

[ ] × [ ] × [ ] × [ ]

4. Solvay Darts: Write the number from the dart into the appropriate space on the dartboard following.

1 substance used for glass and soap manufacture worldwide

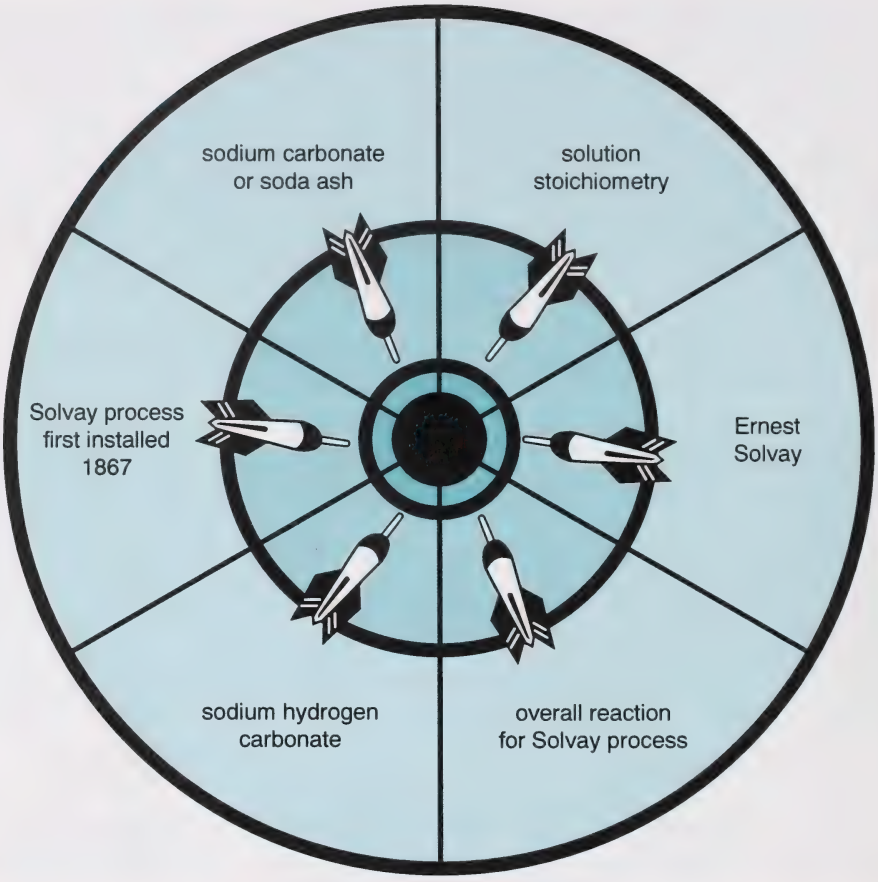
2 $CaCO_{3(s)} + 2\,NaCl_{(aq)} \rightarrow Na_2CO_{3(aq)} + CaCl_{2(aq)}$

3 Belgium chemist who discovered process through experimentation

4 separates from ammonium chloride in cold water by precipitation

5 helps calculate volumes of solutions necessary in the Solvay process

6 lower cost and much less air pollution than the older methods of production



Check your answers by turning to the Appendix, Section 2: Extra Help.

Enrichment

Do any or all of the following.

1. In cellular respiration, glucose $(C_6H_{12}O_6)$ in the cells is broken down when reacted with oxygen to form carbon dioxide and water. If 1.00 g of glucose were completely used, what volume of carbon dioxide at SATP would be released?

2. Both Haber and Bosch received Nobel prizes for their contributions to chemistry. Write a report on the scientific discoveries of Haber and Bosch and indicate the value of their process to modern chemical technology.

3. Some chemical compounds are said to be nonstoichiometric in their atomic ratios. Relate how scientists are able to explain nonstoichiometric chemicals to chemical concepts. Discuss two examples of nonstoichiometric chemical compounds.

Check your answers by turning to the Appendix, Section 2: Enrichment.

Conclusion

In this section, you learned how to deal with gas and solution stoichiometry. This gave you a considerably wider range of problem-solving strategies when combined with your prior knowledge of gravimetric stoichiometry. With your knowledge of stoichiometry, you considered several examples from industrial processes. Besides a brief look at the Haber process and the Down's Cell, you studied the Solvay process in some depth.

50

Section 2 Assignment: Applications of Stoichiometry

Review the Evaluation information found in the introductory pages of this module.

It is important to number and clearly identify each page with the following information at the top:

Chemistry 20 – Module 3 Section 2 Assignment Page # Name and ID #

Be sure to write legibly. Leave a wide left margin and number all of your pages.

1. The conversion factors on the left, when arranged properly, will convert the values on the right to moles. Match the conversion factors to the corresponding values. **(4 marks)**

a. 24.8 L/mol	A. 58.23 mL $O_{2(g)}$ at STP
b. 22.99 g/mol	B. 3.691 g $NaOH_{(s)}$
c. 22.4 L/mol	C. 18 kL $NH_{3(g)}$ at 450 kPa and 80°C
d. $\dfrac{8.31\dfrac{\text{kPa}\cdot\text{L}}{\text{mol}\cdot\text{K}}\times 353\text{ K}}{450\text{ kPa}}$	D. 15.8 mL $H_{2(g)}$ at SATP

2. In the stoichiometric method, what two steps are common to all processes, no matter whether you are working in gravimetric, gas, or solution stoichiometry? **(2 marks)**

3. You are given the following stoichiometry problem:

   What mass of water would be produced in the complete combustion of hydrogen gas with 25.2 L of oxygen gas measured at 23.2°C and 125 kPa?

   a. Describe in words the steps you would perform to solve the problem. **(4 marks)**

   b. Solve the given problem. Show your work. **(5 marks)**

4. Octane, $C_8H_{18(l)}$, mixes with air (oxygen) in the engine of a car and burns with explosive force. For each 1.000 g of octane that completely burns, what volume of carbon dioxide gas measured at SATP would be produced? Show your work. **(5 marks)**

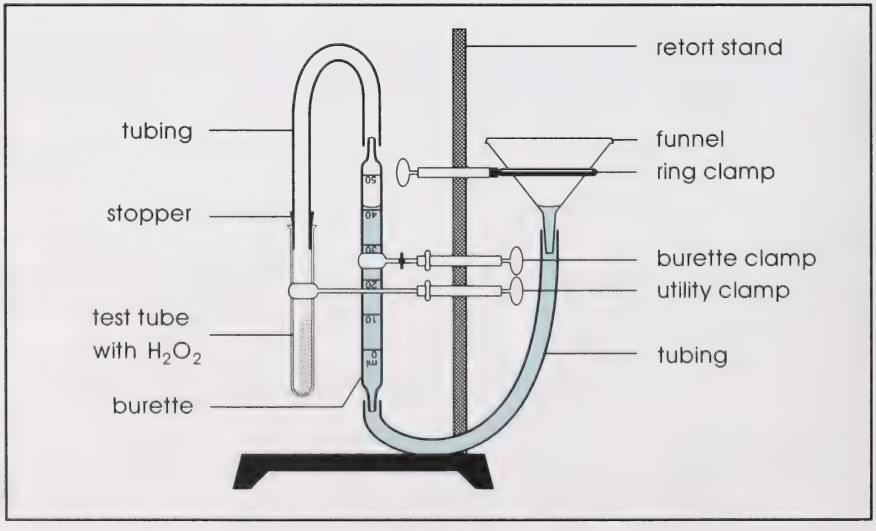
5. Calculate the volume of 0.020 mol/L hydrochloric acid that can be completely neutralized by reaction with 0.920 g of aluminum hydroxide in an antacid tablet. Show your work. **(5 marks)**

6. When copper metal is placed in a solution of nitric acid, a toxic brown gas, nitrogen dioxide, is produced along with aqueous copper(II) nitrate and water. What volume of nitrogen dioxide gas, measured at 20.0°C and 108 kPa, would be produced from the complete reaction of 95.2 mL of 0.115 mol/L nitric acid with copper? Show your work. **(5 marks)**

Science Skills

- ☐ A. Initiating
- ☐ B. Collecting
- ☑ C. Organizing
- ☑ D. Analysing
- ☐ E. Synthesizing
- ☑ F. Evaluating

7. Three students working on a project designed an experiment to collect the oxygen gas which is produced when hydrogen peroxide, H_2O_2, decomposes. The other product formed in the reaction is water. Their experimental design included a setup as in the following diagram.



The following procedure was followed:

Take the stopper out of the test tube. Using the funnel, fill the burette to a level of 42.8 mL with cold tap water. Add 28.0 mL of 0.0880 mol/L $H_2O_{2(aq)}$ to the test tube. Stopper the test tube. Leave the setup to stand for thirty minutes prior to starting the reaction.

The following observations were made:

The room temperature was measured at 19.0°C and the weather station said that the air pressure was 101.2 kPa.

The reaction was started by a catalyst (a substance that speeds up the process, but has no effect on the overall reaction) and allowed to go to completion. The oxygen was collected in the burette. When the two water levels in the burette and funnel were equalized, the final burette reading was 15.0 mL.

a. Complete a prediction for this investigation. Show your work. **(4 marks)**

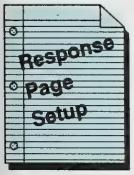
b. Organize the evidence in a table. **(2 marks)**

c. Why did the students wait for thirty minutes after setting up the apparatus? **(1 mark)**

d. Why did the students lower the funnel to equalize the water levels in the burette and funnel? **(2 marks)**

e. Do a brief analysis of the data. **(1 mark)**

f. Evaluate the experiment. Include a percent difference calculation. List two possible sources of error. **(4 marks)**



Use the following heading to organize your response.

- % difference
- Evaluation
- Sources of Error

Reproduce the science skills assessment box after your response to this question. Remember to indicate your evaluation of your skill level for each identified science skill. Be sure to include the spaces for teacher assessment.

Self:	A.	☐	B.	☐	C.	☐	D.	☐	E.	☐	F.	☐
Teacher:	A.	☐	B.	☐	C.	☐	D.	☐	E.	☐	F.	☐

8. You have studied several examples where stoichiometry has been applied to industrial processes.

   Choose one of the processes. In a paragraph (1/2 to 3/4 page) describe the process briefly, highlighting important aspects. Write the overall reaction representing the process. Indicate one situation where stoichiometry could be used in the process to predict outcomes and solve a problem. Also indicate what factors would influence how much product is produced. **(6 marks)**

MODULE SUMMARY

In this module you reviewed and expanded your knowledge of chemical equations and the mole concept. When you combine this knowledge with mole/mass calculations, you are able to solve stoichiometry problems. These calculations help you predict quantities of reactants and products in chemical reactions.

You discovered three related types of stoichiometry: gravimetric, gas, and solution stoichiometry. Regardless of the type of stoichiometry, or in fact if the problems were of a mixed variety, you learned that all stoichiometry problems have common steps and methods. The bridge between given and required substances is always the mole ratio, which comes from a balanced chemical equation.

Testing the validity of designs and methods is always an important scientific process. You performed and analysed several investigations which enabled you to validate the stoichiometric method.

It is always important to see new-found knowledge put to practical use. Your analysis of several industrial processes gave you a glimpse of how stoichiometric methods can be used in industry and society.

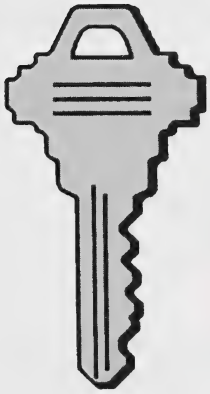
To ensure that all of your work has been completed in a satisfactory manner, check off the items in the following list:

- ❒ Section 1 Assignment has been completed.
- ❒ Section 2 Assignment has been completed.
- ❒ Your responses are oganized and neat, with room for teacher comments.
- ❒ All your response pages are numbered consecutively and identified with this heading.

Chemistry 20 – Module 3 Section # Assignment Page # Name and ID#

Submit **only** your **assignment response pages** for evaluation.

Appendix



Glossary

Activities

Extra Help

Enrichment

Glossary

assumption: the acceptance of an idea without proof of its truth

Avogadro's constant (Avogadro's number): the number of entities or particles $\left(6.02 \times 10^{23}\right)$ in one mole of a substance

chemical equation: the formulas of the reactants on the left and the formulas of the products represented on the right connected with an arrow representing a chemical reaction

chemical formulas: a shorthand way of showing the number and type of atoms present in a substance

Factor Label method: a method of solving all the steps of a problem (e.g., stoichiometry) by using units arranged on one line so that the units can be cancelled, revealing the final unit of the solution

gravimetric stoichiometry: stoichiometry dealing with mass and mole quantities

molar mass: the number of grams of a chemical substance in one mole $\left(6.02 \times 10^{23} \text{ particles}\right)$ of that substance

molar volume: the number of litres in volume occupied by one mole of a gas

mole: the amount of a substance that contains 6.02×10^{23} entities (particles)

mole ratio: a whole number ratio between two substances involved in a chemical reaction

percent difference: the difference between the predicted value and the experimental or actual value, expressed as a percentage of the predicted value

relative atomic mass: the mass of one chemical compared to another (standard mass – Carbon-12) when equal numbers of atoms are used

relative mass: the mass of one substance compared to another when equal numbers of both substances are involved

stoichiometry: the use of proportions to calculate quantities of substances involved in chemical reactions

Suggested Answers

Section 1: Activity 1

1. Chemical reactions are represented on paper by balanced chemical equations.

2. Without the correct formulas, balancing the equation is next to impossible. Even if balancing were possible, the mole ratios would be incorrect.

3. Textbook question 2:

The primary perspective is scientific.

$2\ SO_{2(g)} + O_{2(g)} \rightarrow 2\ SO_{3(g)}$

Textbook question 3:

The primary perspective is political.

$SO_{3(g)} + H_2O_{(l)} \rightarrow H_2SO_{4(aq)}$

Textbook question 4:

The primary perspective is technological.

$2\ CaO_{(s)} + 2\ SO_{2(g)} + O_{2(g)} \rightarrow 2\ CaSO_{4(s)}$

Textbook question 5:

The primary perspective is economic.

$CaO_{(s)} + H_2SO_{3(aq)} \rightarrow H_2O_{(l)} + CaSO_{3(s)}$

Textbook question 6:

The primary perspective is ecological.

$Al_2(SiO_3)_{3(s)} + 3\ H_2SO_{4(aq)} \rightarrow 3\ H_2SiO_{3(aq)} + Al_2(SO_4)_{3(aq)}$

4.

Reaction Type	General Description
formation	elements → **compound**
simple decomposition	**compound** → elements
complete combustion	substance + oxygen → **common oxides**
single replacement	**element + compound** → element + compound
double replacement	**compound + compound → compound + compound**

5. Textbook question 7:

Textbook Question	Reaction Type	Balanced Equation
a	formation	$2\ Al_{(s)} + 3\ F_{2(g)} \rightarrow 2\ AlF_{3(s)}$
b	simple decomposition	$2\ NaCl_{(s)} \rightarrow 2\ Na_{(s)} + Cl_{2(g)}$
c	formation	$S_{8(g)} + 8\ O_{2(g)} \rightarrow 8\ SO_{2(g)}$
d	complete combustion	$CH_{4(g)} + 2\ O_{2(g)} \rightarrow CO_{2(g)} + 2\ H_2O_{(g)}$
e	simple decomposition	$2\ Al_2O_{3(s)} \rightarrow 4\ Al_{(s)} + 3\ O_{2(g)}$
f	complete combustion	$C_3H_{8(g)} + 5\ O_{2(g)} \rightarrow 3\ CO_{2(g)} + 4\ H_2O_{(g)}$
g	formation	$2\ Hg_{(l)} + O_{2(g)} \rightarrow 2\ HgO_{(s)}$
h	simple decomposition	$2\ FeBr_{3(s)} \rightarrow 2\ Fe_{(s)} + 3\ Br_{2(l)}$
i	complete combustion	$2\ C_4H_{10(g)} + 13\ O_{2(g)} \rightarrow 8\ CO_{2(g)} + 10\ H_2O_{(g)}$

6. a. v c. iv e. iii
   b. i d. vi f. ii

7. a. $Na_2SO_{4(aq)}$ c. $CaCO_{3(s)}$ e. $Al(OH)_{3(s)}$
   b. $CuBr_{2(aq)}$ d. $K_2SO_{3(aq)}$

8. Single and double replacement reactions usually occur in the presence of water (in aqueous solution). The other reactions usually involve substances in their pure states – solid, liquid, or gas.

9. Textbook question 14:

   a. The reaction is simple decomposition. $2\ NaCl_{(s)} \rightarrow 2\ Na_{(s)} + Cl_{2(g)}$

   b. The reaction is formation. $4\ Na_{(s)} + O_{2(g)} \rightarrow 2\ Na_2O_{(s)}$

   c. The reaction is single replacement. $2\ Na_{(s)} + 2\ HOH_{(l)} \rightarrow H_{2(g)} + 2\ NaOH_{(aq)}$

   d. The reaction is double replacement. $AlCl_{3(aq)} + 3\ NaOH_{(aq)} \rightarrow Al(OH)_{3(s)} + 3\ NaCl_{(aq)}$

   e. The reaction is single replacement. $2\ Al_{(s)} + 3\ H_2SO_{4(aq)} \rightarrow 3\ H_{2(g)} + Al_2(SO_4)_{3(aq)}$

   f. The reaction is complete combustion. $2\ C_8H_{18(l)} + 25\ O_{2(g)} \rightarrow 16\ CO_{2(g)} + 18\ H_2O_{(g)}$

10. Textbook question 8:

    a. The reaction is a single replacement . $Br_{2(l)} + 2\ NaI_{(aq)} \rightarrow I_{2(s)} + 2\ NaBr_{(aq)}$

    b. The reaction is a double replacement . $H_2SO_{4(aq)} + 2\ NaOH_{(aq)} \rightarrow 2\ H_2O_{(l)} + Na_2SO_{4(aq)}$

11. The term relative mass is used when masses of different types of atoms are compared to a chosen standard.

12. Chemists use the term *relative mass* because they cannot mass and compare single atoms of each kind of atom. They must take the total mass of a large number of one kind of atom and then compare the total mass to the same number of another kind of atom.

13. A container with the same number of atoms is desirable for the purpose of comparing different substances.

14. You can compare any number of units of two substances, as long as the number of units is the same for both substances. The relative mass ratio would be the same.

15. a. Oxygen was the first element used as a standard.

    b. Carbon-12, the isotope of carbon containing six protons and six neutrons, is the element used as the standard today.

16. The mass spectrometer allows scientists to find the relative mass of atoms of any substance.

17. Since gases that react are in the form of molecules, the container could not be applied to elements that exist naturally in the form of atoms.

18. An Aardvark is a hypothetical term specific to this video. It would be too difficult to count the number of items required in an Aardvark $\left(1 \times 10^{6} \text{ items}\right)$.

19. Scientists picked Avogadro's number -6.02×10^{23} – as a useful number of atoms.

20. Twelve grams of Carbon-12 contain 6.02×10^{23} or Avogadro's number of atoms of carbon.

21. To determine the number of atoms from the relative mass of an element, look up the relative mass of any element, and measure out that mass in grams. This amount always contains the same number of atoms of that element, namely Avogadro's number of atoms.

22. Avogadro's number of atoms is called a mole.

23. One mole or Avogadro's number of particles of a gas occupies 22.4 L at 101.3 kPa and 273 K (STP), or 24.8 L at 100 kPa and 298 K (STP).

24. There are about 6.023×10^{23} atoms in a mole.

25. Using the mole is a practical and universal way of knowing the same number of atoms. The mass can be easily determined in the laboratory and you can compare the mass of individual atoms and molecules as they combine and recombine in chemical reactions.

26. The SI symbol for Avogadro's constant is N_A.

27. A balanced chemical equation can be read in terms of molecules or moles because the numbers are in the same relative ratio, no matter what number unit is used.

28. Textbook question 1:

    a. $2\ Al_{(s)} + 3\ CuSO_{4(aq)} \rightarrow 3\ Cu_{(s)} + Al_2(SO_4)_{3(aq)}$

    b. $2 \text{ mol } Al : 3 \text{ mol } CuSO_4 : 3 \text{ mol } Cu : 1 \text{ mol } Al_2(SO_4)_3$

Section 1: Activity 2

1. Since the combustion reaction has a ratio of $2\ H_2 : 1\ O_2$, the shuttle needs $2\left(2 \times 10^{7} \text{ mol}\right)$ or 4×10^{7} mol of liquid hydrogen.

2. Textbook question 15:

The molar mass of the various gases are as follows:

a. $Hg = 200.59\ g / mol$

b. $O_3 = 48.00\ g / mol$

c. $CO_2 = 44.01\ g / mol$

d. $NH_3 = 17.04\ g / mol$

e. $H_2S = 34.08\ g / mol$

f. $N_2O_4 = 92.02\ g / mol$

g. $ClO_2 = 67.45\ g / mol$

Textbook question 16:

Molar mass gave chemists a universal container to compare relative masses of different substances.

Textbook question 17:

a. $n_{(C_8H_{18})} = 50\ \cancel{g\ C_8H_{18}} \times \dfrac{1\ mol}{114.26\ \cancel{g\ C_8H_{18}}}$

$= 0.44\ mol$

b. $n_{(C_8H_{18})} = 70.0\ \cancel{g\ CH_3OH} \times \dfrac{1\ mol}{32.05\ \cancel{g\ CH_3OH}}$

$= 2.18\ mol$

c. $500\ mg\ Cl_2 = 0.500\ g\ Cl_2$

$n_{(Cl_2)} = 0.500\ \cancel{g\ Cl_2} \times \dfrac{1\ mol}{70.90\ \cancel{g\ Cl_2}}$

$= 0.00705\ mol$

or

$7.05 \times 10^{-3}\ mol$

Textbook question 18:

a. $m_{(CCl_2F_2)} = 2.50\ \cancel{mol\ CCl_2F_2} \times \dfrac{120.91\ g}{1\ \cancel{mol\ CCl_2F_2}}$

$= 302\ g$

b. $15\ mmol\ Rn = 0.015\ mol\ Rn$

$m_{(Rn)} = 0.015\ \cancel{mol\ Rn} \times \dfrac{222\ g}{1\ \cancel{mol\ Rn}}$

$= 3.3\ g$

or

$15\ m\cancel{mol\ Rn} \times \dfrac{222\ g}{1\ \cancel{mol\ Rn}}$

$= 3.3 \times 10^3\ mg$

c. $20.0\ kmol\ O_2 = 2.00 \times 10^4\ mol\ O_2$

$m_{(O_2)} = 2.00 \times 10^4\ \cancel{mol\ O_2} \times \dfrac{32.00\ g}{1\ \cancel{mol\ O_2}}$

$= 6.40 \times 10^5\ g$

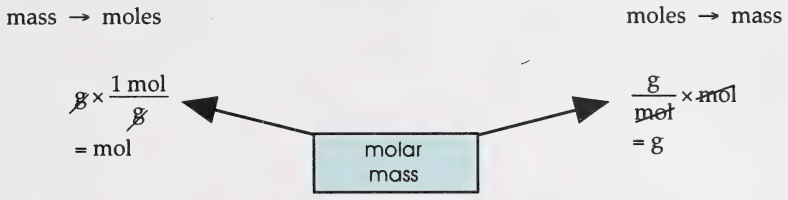
or

$20.0\ k\cancel{mol\ O_2} \times \dfrac{32.00\ g}{1\ \cancel{mol\ O_2}}$

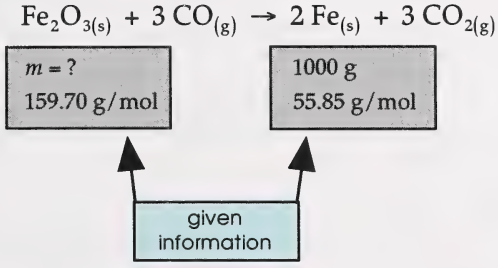
$= 640\ kg$

3. The coefficients indicate the relative mole ratio of reactants and products in a chemical reaction.

4. The molar mass allows for conversions between mass and amount (moles).



5. The given information is written directly under the appropriate substances in the chemical equation.



6. **Step 1:** **Write a balanced chemical equation.**

CALCULATION

Step 2	Step 3	Step 4
Convert mass given to moles given.	Convert moles given to moles required.	Convert moles required to mass required.

Given Information → Required Information

In step 3 (mole ratio), you convert **given** information about a substance to information you **require** about another substance.

7. Textbook question 9:

Step 1

$8\ Zn_{(s)} + S_{8(s)} \rightarrow 8\ ZnS_{(s)}$

25 g — m = ?

65.38 g/mol — 97.44 g/mol

Step 2

$n_{(Zn)} = 25\ \not{g\ Zn} \times \frac{1\ \text{mol Zn}}{65.38\ \not{g\ Zn}}$

$= 0.38\ \text{mol Zn}$

Step 3

$n_{(S_8)} = 0.38\ \not{\text{mol Zn}} \times \frac{1\ \text{mol } S_8}{8\ \not{\text{mol Zn}}}$

$= 0.048\ \text{mol } S_8$

Step 4

$m_{(S_8)} = 0.048\ \not{\text{mol } S_8} \times \frac{97.44\ \text{g}}{1\ \not{\text{mol } S_8}}$

$= 4.7\ \text{g}$

You would react 4.7 g of sulphur with 25 g of zinc.

8. Textbook question 10:

$2\,Al_2O_{3(s)} \rightarrow 4\,Al_{(s)} + 3\,O_{2(g)}$

100 g $\quad m = ?$

101.96 g/mol $\quad 26.98$ g/mol

$$m_{(Al)} = 100\ \text{g}\ \cancel{Al_2O_3} \times \frac{1\ \cancel{\text{mol}\ Al_2O_3}}{101.96\ \cancel{\text{g}\ Al_2O_3}} \times \frac{4\ \cancel{\text{mol Al}}}{2\ \cancel{\text{mol}\ Al_2O_3}} \times \frac{26.98\ \text{g}}{1\ \cancel{\text{mol Al}}}$$

$$= 52.9\ \text{g}$$

You would produce 52.9 g of aluminum in the reaction.

Textbook question 11:

$C_3H_{8(g)} + 5\,O_{2(g)} \rightarrow 3\,CO_{2(g)} + 4\,H_2O_{(g)}$

10.0 g $\quad m = ?$

44.11 g/mol $\quad 32.00$ g/mol

$$m_{(O_2)} = 10.0\ \cancel{\text{g}\ C_3H_8} \times \frac{1\ \cancel{\text{mol}\ C_3H_8}}{44.11\ \cancel{\text{g}\ C_3H_8}} \times \frac{5\ \cancel{\text{mol}\ O_2}}{1\ \cancel{\text{mol}\ C_3H_8}} \times \frac{32.00\ \text{g}}{1\ \cancel{\text{mol}\ O_2}}$$

$$= 36.3\ \text{g}$$

It would take 36.3 g of oxygen to completely react the propane.

Textbook question 12:

$2\,NaCl_{(aq)} + Pb(NO_3)_{2(aq)} \rightarrow 2\,NaNO_{3(aq)} + PbCl_{2(s)}$

2.57 g $\quad m = ?$

58.44 g/mol $\quad 278.10$ g/mol

$$m_{(PbCl_2)} = 2.57\ \cancel{\text{g NaCl}} \times \frac{1\ \cancel{\text{mol NaCl}}}{58.44\ \cancel{\text{g NaCl}}} \times \frac{1\ \cancel{\text{mol}\ PbCl_2}}{2\ \cancel{\text{mol NaCl}}} \times \frac{278.10\ \text{g}}{1\ \cancel{\text{mol}\ PbCl_2}}$$

$$= 6.11\ \text{g}$$

The reaction would produce 6.11 g of lead(II) chloride precipitate.

Textbook question 13:

$2\,Al_{(s)} + 3\,H_2SO_{4(aq)} \rightarrow 3\,H_{2(g)} + Al_2(SO_4)_{3(aq)}$

2.73 g $\quad m = ?$

26.98 g/mol $\quad 2.02$ g/mol

$$m_{(H_2)} = 2.73\ \cancel{g\ Al} \times \frac{1\ \cancel{mol\ Al}}{26.98\ \cancel{g\ Al}} \times \frac{3\ \cancel{mol\ H_2}}{2\ \cancel{mol\ Al}} \times \frac{2.02\ g}{1\ \cancel{mol\ H_2}}$$

$$= 0.307\ g$$

The reaction would produce 0.307 g of hydrogen gas.

Textbook question 14:

$$2\,KOH_{(aq)} + Cu(NO_3)_{2(aq)} \rightarrow 2\,KNO_{3(aq)} + Cu(OH)_{2(s)}$$

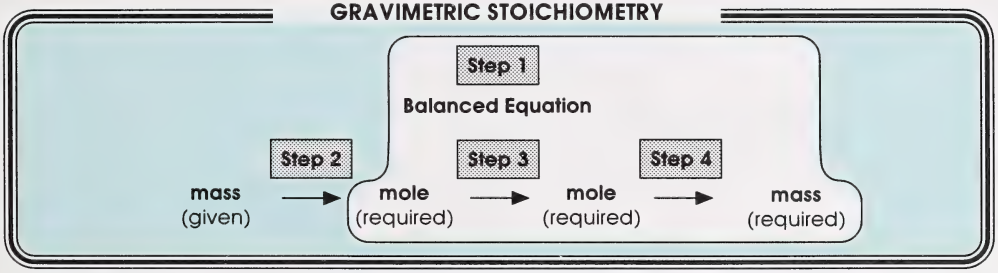
2.67 g	m = ?
56.11 g/mol	97.57 g/mol

$$m_{(Cu(OH)_2)} = 2.67\ \cancel{g\ KOH} \times \frac{1\ \cancel{mol\ KOH}}{56.11\ \cancel{g\ KOH}} \times \frac{1\ \cancel{mol\ Cu(OH)_2}}{2\ \cancel{mol\ KOH}} \times \frac{97.57\ g}{1\ \cancel{mol\ Cu(OH)_2}}$$

$$= 2.32\ g$$

You produce 2.32 g of copper(II) hydroxide precipitate in the reaction.

9.



10. There are many possible answers. However, the question must contain the following:

- a description of the reaction
- a given mass of one of the reactants or products
- a question asking for moles of one of the substances – a different substance than the "given mass" chemical

Here is one possible example: What number of moles of oxygen is required to completely react with 2.52 g of magnesium metal to produce a white ionic solid?

11. All the stoichiometry problems need a balanced chemical equation (Step 1) and a mole ratio calculation (Step 3).

12. No, the coefficients are only used in Step 3 – the mole ratio calculation. The molar mass conversion factor is always the mass per **one** mole.

13. $CH_{4(g)} + 2\,O_{2(g)} \rightarrow CO_{2(g)} + 2\,H_2O_{(g)}$

0.951 mol $\quad n = ?$

$$n_{(O_2)} = 0.951 \text{ mol } CH_4 \times \frac{2 \text{ mol } O_2}{1 \text{ mol } CH_4}$$

$$= 1.90 \text{ mol } O_2$$

You use 1.90 mole O_2 in the reaction.

14. $N_{2(g)} + O_{2(g)} \rightarrow 2\,NO_{(g)}$

$m = ?$ $\quad$ 15.07 mol

28.02 g/mol

$$m_{(N_2)} = 15.07 \text{ mol NO} \times \frac{1 \text{ mol } N_2}{2 \text{ mol NO}} \times \frac{28.02 \text{ g}}{1 \text{ mol } N_2}$$

$$= 211.1 \text{ g}$$

The mass of nitrogen that reacts is 211.1 g.

15. $6\,Li_{(s)} + N_{2(g)} \rightarrow 2\,Li_3N_{(s)}$

7.36 g $\quad n = ?$

6.94 g/mol

$$n_{(Li_3N)} = 7.36 \text{ g Li} \times \frac{1 \text{ mol Li}}{6.94 \text{ g Li}} \times \frac{2 \text{ mol } Li_3N}{6 \text{ mol Li}}$$

$$= 0.354 \text{ mol } Li_3N$$

You produce 0.354 mol Li_3N.

16. $2\,Al_{(s)} + Fe_2O_{3(s)} \rightarrow Al_2O_{3(s)} + 2\,Fe_s$

25.0 g $\quad m = ?$

26.98 g/mol $\quad$ 159.70 g/mol

$$m_{(Fe_2O_3)} = 25.0 \text{ g Al} \times \frac{1 \text{ mol Al}}{26.98 \text{ g Al}} \times \frac{1 \text{ mol } Fe_2O_3}{2 \text{ mol Al}} \times \frac{159.70 \text{ g}}{1 \text{ mol } Fe_2O_3}$$

$$= 74.0 \text{ g}$$

The mass of $Fe_2O_{3(s)}$ required for every 25.0 g of aluminum used is 74.0 g.

17. The most rigorous test of any scientific concept is whether or not you can use it to make predictions.

18. The prediction is falsified if the percent difference between the actual and the predicted values is greater than 10%.

19. a. A common experimental design for testing stoichiometric predictions is to use precipitation filtration techniques.

    b. Stoichiometry is used to predict the possible outcome of the investigation; the experimental design and process tests the actual outcome. This confirms or refutes the prediction.

20. The vertical lines represent the absolute value of the number or operation. This means that whether the values are positive or negative, you work with only positive outcomes.

21. The $\times 100$ turns the value into a percentage.

22. The prediction is valid for any value less than 10% difference.

23. $Zn_{(s)} + Pb(NO_3)_{2(aq)} \rightarrow Zn(NO_3)_{2(aq)} + Pb_{(s)}$

2.13 g	$m = ?$
65.38 g/mol	207.20 g/mol

24. $$m_{(Pb)} = 2.13 \text{ g Zn} \times \frac{1 \text{ mol Zn}}{65.38 \text{ g Zn}} \times \frac{1 \text{ mol Pb}}{1 \text{ mol Zn}} \times \frac{207.20 \text{ g}}{1 \text{ mol Pb}}$$
$$= 6.75 \text{ g}$$

    The amount of lead predicted to be produced is 6.75 g.

25. The mass of the filter paper must be subtracted from the total mass (the lead plus the filter paper) to get the mass of the lead produced.

26. The mass of lead and filter paper is 7.60 g.
    The mass of filter paper is 0.92 g.
    Therefore the mass of lead is 6.68 g (7.60 g – 0.92 g = 6.68 g).

27. a. The experimental design was judged to be adequate because the problem was answered without any obvious flaws. The design seems to be the most efficient available.

    An evaluation of the prediction is as follows:

$$\% \text{ difference} = \frac{|\text{experimental} - \text{predicted}|}{\text{predicted}} \times 100$$
$$= \frac{|6.68 \text{ g} - 6.75 \text{ g}|}{6.75 \text{ g}} \times 100$$
$$= 1.04\%$$

    The prediction is judged to be verified because the percent difference is very small – significantly less than 10%.

    b. The prediction was valid because the percent difference was less than 10%.

28. The procedure required for making the prediction is gravimetric stoichiometry.

29. a. The precipitate is $Ca_3(PO_4)_{2(s)}$ – calcium phosphate.

b. $2\ Na_3PO_{4(aq)} + 3\ Ca(NO_3)_{2(aq)} \rightarrow 6\ NaNO_{3(aq)} + Ca_3(PO_4)_{2(s)}$

2.98 g	m = ?
163.94 g/mol	310.18 g/mol

$$m_{(Ca_3(PO_4)_2)} = 2.98\ \cancel{g\ Na_3PO_4} \times \frac{1\ \cancel{mol\ Na_3PO_4}}{163.94\ \cancel{g\ Na_3PO_4}} \times \frac{1\ \cancel{mol\ Ca_3(PO_4)_2}}{2\ \cancel{mol\ Na_3PO_4}} \times \frac{310.18\ g}{1\ \cancel{mol\ Ca_3(PO_4)_2}}$$

$$= 2.82\ g$$

The amount of precipitate formed is 2.82 g $Ca_3(PO_4)_2$.

30. A known mass of sodium phosphate is mixed and reacted with an excess of calcium nitrate. The precipitate formed is filtered, dried, and massed.

31. The materials required in the experiment are as follows:

- calcium nitrate solution
- distilled water
- funnel
- retort stand
- stirring rod
- watch glass
- sodium phosphate
- filter paper
- beakers
- funnel clamp (or normal clamp)
- balance

To see a photograph of the filtration apparatus, turn to page 534 of *Nelson Chemistry* (Figure C15).

32. The procedure is as follows:

- Measure 2.98 g of sodium phosphate in a beaker on a balance.
- Add enough distilled water to the sodium phosphate to dissolve the solid.
- Measure 4.60 g* of calcium nitrate in a beaker on a balance.

* The mass of calcium nitrate needed can be found by using stoichiometry.
What mass of calcium nitrate is needed to react with 2.98 g of $Na_3(PO_4)_2$?

$2\ Na_3PO_{4(aq)} + 3\ Ca(NO_3)_{2(aq)} \rightarrow 6\ NaNO_{3(aq)} + Ca_3(PO_4)_{2(s)}$

2.98 g	m = ?
163.94 g/mol	164.10 g/mol

$$m_{(Ca(NO_3)_2)} = 2.98\ \cancel{g\ Na_3PO_4} \times \frac{1\ \cancel{mol\ Na_3PO_4}}{163.94\ \cancel{g\ Na_3PO_4}} \times \frac{3\ \cancel{mol\ Ca(NO_3)_2}}{2\ \cancel{mol\ Na_3PO_4}} \times \frac{164.10\ g}{1\ \cancel{mol\ Ca(NO_3)_2}}$$

$$= 4.47\ g$$

Since excess $Ca(NO_3)_2$ is needed, use 4.60 g $Ca(NO_3)_{2(s)}$.

- Dissolve the $Ca(NO_3)_{2(s)}$ in distilled water.
- Combine the two solutions in one beaker.
- Set up the filtration apparatus, and follow the filtration procedure, as described on page 534 of your textbook.

33. The analysis of the investigation is as follows:

mass of filter paper + precipitate =	3.82 g
– mass of filter paper =	0.93 g
mass of precipitate =	2.89 g

According to the evidence gathered, the mass of precipitate formed is 2.89 g.

34. You need to address each component of the evaluation including comparing the predicted answer with the experimental answer, checking the authority with which the prediction was made, and deciding if the technological skills were adequate to answer the question. You might even suggest ways to improve the experimental design.

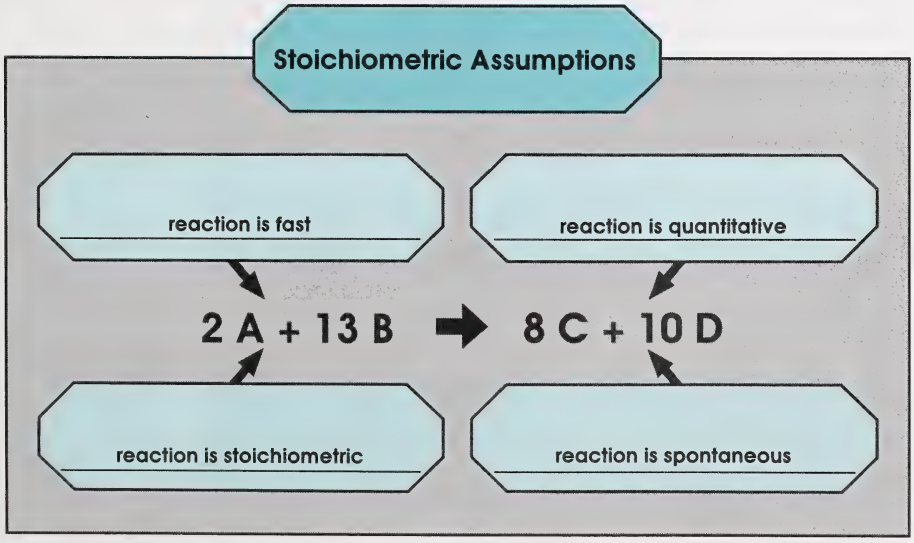
Example answer: $\%\ \text{difference} = \dfrac{|2.89\ \text{g} - 2.82\ \text{g}|}{2.82\ \text{g}} \times 100$

$= 2.5\%$

Since the percent difference is only about 2.5%, the predictions made by stoichiometry must be judged to be adequate. The techniques are adequate to measure the precipitate to an accuracy of 0.01 g. Calculations are also to an accuracy of 0.01 g. The experiment could be improved by checking the filtrate to ensure that all the sodium phosphate reacted.

Section 1: Activity 3

1.



2. $Sr(NO_3)_{2(aq)} + CuSO_{4(aq)} \rightarrow SrSO_{4(s)} + Cu(NO_3)_{2(aq)}$

2.00 g	m = ?
211.64 g/mol	183.68 g/mol

$$m_{(SrSO_4)} = 2.00 \text{ g } \cancel{Sr(NO_3)_2} \times \frac{1 \text{ } \cancel{\text{mol } Sr(NO_3)_2}}{211.64 \text{ g } \cancel{Sr(NO_3)_2}} \times \frac{1 \text{ } \cancel{\text{mol } SrSO_4}}{1 \text{ } \cancel{\text{mol } Sr(NO_3)_2}} \times \frac{183.68 \text{ g}}{1 \text{ } \cancel{\text{mol } SrSO_4}}$$

$$= 1.74 \text{ g}$$

The amount of precipitate formed is predicted to be 1.74 g $SrSO_4$.

3. A known mass of strontium nitrate is placed in a beaker with an excess of copper(II) sulphate. Strontium sulphate which precipitates from the reaction is separated by filtration and dried. The mass of the strontium sulphate is determined.

4. The materials required for this investigation are as follows:

- $Sr(NO_3)_{2(s)}$
- distilled water
- laboratory scoop
- filter paper
- beakers
- $CuSO_4 \cdot 5H_2O_{(s)}$
- pan balance
- stirring rod
- retort stand and clamp
- funnel

5. The procedure for the investigation is as follows:

- Measure 2.00 g of $Sr(NO_3)_{2(s)}$ in a clean, dry 250 mL beaker on a balance.
- Add 50 mL distilled water to dissolve the solid.
- Dissolve a minimum of 2.56 g $CuSO_4 \cdot 5H_2O$ into 50 mL of distilled water.
- Slowly add the copper(II) sulphate solution to the strontium nitrate solution.
- Wait for the precipitate to settle, forming a clear layer at the top of the beaker.
- If cloudiness remains, add additional copper(II) sulphate solution until no further precipitation occurs.
- Follow the filtration procedure given on page 534 of your textbook, *Nelson Chemistry*.

6. The evidence may be shown as follows:

Mass of filter paper	1.01 g
Mass of filter paper plus precipitate	2.62 g

7. The analysis may be shown as follows:

mass of filter paper plus precipitate	= 2.62 g
– mass of filter paper	= 1.01 g
mass of precipitate	= 1.61 g

8. The evaluation of the investigation may be as follows:

$$\% \text{ difference} = \frac{|1.61\text{ g} - 1.74\text{ g}|}{1.74\text{ g}} \times 100$$
$$= 7.47\%$$

The experimental design is consistent with the design used to test gravimetric stoichiometry used in the textbook. This design tested the prediction. The sources of error include the uncertainty of the measurements and the dryness of the precipitate. The percent difference at 7.47% is higher than the examples in the textbook, but still less than 10%. The prediction made by the method of gravimetric stoichiometry is verified.

9. $Sr(NO_3)_{2(aq)} + CuSO_{4(aq)} \rightarrow SrSO_{4(s)} + Cu(NO_3)_{2(aq)}$

2.00 g	m = ?
211.64 g/mol	183.68 g/mol

$$m_{(SrSO_4)} = 2.00\text{ g } \cancel{Sr(NO_3)_2} \times \frac{1\text{ } \cancel{\text{mol } Sr(NO_3)_2}}{211.64\text{ g } \cancel{Sr(NO_3)_2}} \times \frac{1\text{ } \cancel{\text{mol } SrSO_4}}{1\text{ } \cancel{\text{mol } Sr(NO_3)_2}} \times \frac{183.68\text{ g}}{1\text{ } \cancel{\text{mol } SrSO_4}}$$
$$= 1.74\text{ g}$$

The amount of precipitate formed is predicted to be 1.74 g $SrSO_4$.

10. A known mass of strontium nitrate is placed in a beaker with an excess of copper(II) sulphate. Strontium sulphate which precipitates from the reaction is separated by filtration and dried. The mass of the strontium sulphate is determined.

11. The materials required for this investigation are as follows:

- $Sr(NO_3)_{2(s)}$
- distilled water
- laboratory scoop
- filter paper
- beakers
- $CuSO_4 \cdot 5H_2O_{(s)}$
- pan balance
- stirring rod
- retort stand and clamp
- funnel

12. The procedure for the investigation is as follows:

- Measure 2.00 g of $Sr(NO_3)_{2(s)}$ in a clean, dry 250 mL beaker on a balance.
- Add 50 mL distilled water to dissolve the solid.
- Dissolve a minimum of 2.56 g $CuSO_4 \cdot 5H_2O$ into 50 mL of distilled water.
- Slowly add the copper(II) sulphate solution to the strontium nitrate solution.
- Wait for the precipitate to settle, forming a clear layer at the top of the beaker.
- If cloudiness remains, add additional copper(II) sulphate solution until no further precipitation occurs.
- Follow the filtration procedure given on page 534 of your textbook, *Nelson Chemistry*.

13. The evidence may be shown as follows:

Mass of filter paper	1.01 g
Mass of filter paper plus precipitate	2.62 g

14. The analysis may be shown as follows:

mass of filter paper plus precipitate	= 2.62 g
– mass of filter paper	= 1.01 g
mass of precipitate	= 1.61 g

15. The evaluation of the investigation may be as follows:

$$\% \text{ difference} = \frac{|1.61\text{ g} - 1.74\text{ g}|}{1.74\text{ g}} \times 100$$
$$= 7.47\%$$

The experimental design is consistent with the design used to test gravimetric stoichiometry used in the textbook. This design tested the prediction. The sources of error include the uncertainty of the measurements and the dryness of the precipitate. The percent difference at 7.47% is higher than the examples in the textbook, but still less than 10%. The prediction made by the method of gravimetric stoichiometry is verified.

16. Textbook question 15:

$$Pb(NO_3)_{2(aq)} + 2\,KBr_{(aq)} \rightarrow PbBr_{2(s)} + 2\,KNO_{3(aq)}$$

m = ?	3.65 g
331.22 g/mol	367.00 g/mol

$$m_{(Pb(NO_3)_2)} = 3.65\,\cancel{\text{g PbBr}_2} \times \frac{1\,\cancel{\text{mol PbBr}_2}}{367.00\,\cancel{\text{g PbBr}_2}} \times \frac{1\,\cancel{\text{mol Pb(NO}_3)_2}}{1\,\cancel{\text{mol PbBr}_2}} \times \frac{331.22\text{ g}}{1\,\cancel{\text{mol Pb(NO}_3)_2}}$$
$$= 3.29\text{ g}$$

The mass of lead(II) nitrate in the sample is 3.29 g.

Textbook question 16:

$$2\,AgNO_{3(aq)} + Na_2CrO_{4(aq)} \rightarrow Ag_2CrO_{4(s)} + 2\,NaNO_{3(aq)}$$

m = ?	2.89 g
169.88 g/mol	331.74 g/mol

$$m_{(AgNO_3)} = 2.89\ \cancel{g\ Ag_2CrO_4} \times \frac{1\ \cancel{mol\ Ag_2CrO_4}}{331.74\ \cancel{g\ Ag_2CrO_4}} \times \frac{2\ \cancel{mol\ AgNO_3}}{1\ \cancel{mol\ Ag_2CrO_4}} \times \frac{169.88\ g}{1\ \cancel{mol\ AgNO_3}}$$
$$= 2.96\ g$$

The mass of silver nitrate present was 2.96 g.

Textbook question 17:

$2\,CrCl_{3(aq)} + 3\,Zn_{(s)} \rightarrow 2\,Cr_{(s)} + 3\,ZnCl_{2(aq)}$

m = ?	50.0 g initial
158.35 g/mol	−38.5 g leftover
	11.5 g reacted
	65.38 g/mol

$$m_{(CrCl_3)} = 11.5\ \cancel{g\ Zn} \times \frac{1\ \cancel{mol\ Zn}}{65.38\ \cancel{g\ Zn}} \times \frac{2\ \cancel{mol\ CrCl_3}}{3\ \cancel{mol\ Zn}} \times \frac{158.35\ g}{1\ \cancel{mol\ CrCl_2}}$$
$$= 18.6\ g$$

The mass of chromium(III) chloride in the sample was 18.6 g.

Section 1: Follow-up Activities

Extra Help

1. a. $8\,H_2S_{(g)} \rightarrow 8\,H_{2(g)} + \underset{\uparrow}{1}\,S_{8(g)}$

   It is not necessary to write the 1.

   b. $2\,KI_{(aq)} + Pb(NO_3)_{2(aq)} \rightarrow 2\,KNO_{3(aq)} + PbI_{2(s)}$

   c. $2\,F_{2(g)} + 4\,NaOH_{(aq)} \rightarrow 4\,NaF_{(aq)} + O_{2(g)} + 2\,H_2O_{(l)}$

   d. $Fe(NO_3)_{3(aq)} + 3\,NH_4OH_{(aq)} \rightarrow Fe(OH)_{3(s)} + 3\,NH_4NO_{3(aq)}$

2. a. **Step 3** convert moles of given to moles required

   Step 1 write the balanced chemical equation

   Step 4 convert moles required to mass required

   Step 2 convert mass given to moles given

b. $Cu_{(s)} + 2\ AgNO_{3(aq)} \rightarrow Cu(NO_3)_{2(aq)} + 2\ Ag_{(s)}$

16.2 g — m = ?

63.55 g/mol — 107.87 g/mol

$$m_{(Ag)} = 16.2\ \cancel{g\ Cu} \times \frac{1\ \cancel{mol\ Cu}}{63.55\ \cancel{g\ Cu}} \times \frac{2\ \cancel{mol\ Ag}}{1\ \cancel{mol\ Cu}} \times \frac{107.87\ g}{1\ \cancel{mol\ Ag}}$$

$$= 55.0\ g$$

The mass of silver that can be plated out is 55.0 g.

3. a. Step 1, Step 3, and Step 4 are needed.

b. Step 1, Step 2, and Step 3 are needed.

4. The solution to problem 3. a. is as follows:

$N_{2(g)} + 3\ H_{2(g)} \rightarrow 2\ NH_{3(g)}$

26.8 mol — m = ?

17.04 g/mol

$$m_{(NH_3)} = 26.8\ \cancel{mol\ H_2} \times \frac{2\ \cancel{mol\ NH_3}}{3\ \cancel{mol\ H_2}} \times \frac{17.04\ g}{1\ \cancel{mol\ NH_3}}$$

$$= 304\ g$$

The solution to problem 3. b. is as follows:

$C_3H_{8(g)} + 5\ O_{2(g)} \rightarrow 3\ CO_{2(g)} + 4\ H_2O_{(g)}$

468 g — n = ?

44.11 g/mol

$$n_{(CO_2)} = 468\ \cancel{g\ C_3H_8} \times \frac{1\ \cancel{mol\ C_3H_8}}{44.11\ \cancel{g\ C_3H_8}} \times \frac{3\ mol\ CO_2}{1\ \cancel{mol\ C_3H_8}}$$

$$= 31.8\ mol\ CO_2$$

Enrichment

1. $(NH_4)_2Cr_2O_{7(s)} \rightarrow Cr_2O_{3(s)} + N_{2(g)} + 4\ H_2O_{(g)}$

3.00 g — m = ? — m = ? — m = ?

252.10 g/mol — 152.00 g/mol — 28.02 g/mol — 18.02 g/mol

$$m_{(Cr_2O_3)} = 3.00\ \cancel{g\ (NH_4)_2Cr_2O_7} \times \frac{1\ \cancel{mol\ (NH_4)_2Cr_2O_7}}{252.10\ \cancel{g\ (NH_4)_2Cr_2O_7}} \times \frac{1\ \cancel{mol\ Cr_2O_3}}{1\ \cancel{mol\ (NH_4)_2Cr_2O_7}} \times \frac{152.00\ g}{1\ \cancel{mol\ Cr_2O_3}}$$

$$= 1.81\ g$$

$$m_{(N_2)} = 3.00\ \cancel{g\ (NH_4)_2Cr_2O_7} \times \frac{1\ \cancel{mol\ (NH_4)_2Cr_2O_7}}{252.10\ \cancel{g\ (NH_4)_2Cr_2O_7}} \times \frac{1\ \cancel{mol\ N_2}}{1\ \cancel{mol\ (NH_4)_2Cr_2O_7}} \times \frac{28.02\ g}{1\ \cancel{mol\ N_2}}$$

$$= 0.333\ g$$

$$m_{(H_2O)} = 3.00\ \cancel{g\ (NH_4)_2Cr_2O_7} \times \frac{1\ \cancel{mol\ (NH_4)_2Cr_2O_7}}{252.10\ \cancel{g\ (NH_4)_2Cr_2O_7}} \times \frac{4\ \cancel{mol\ H_2O}}{1\ \cancel{mol\ (NH_4)_2Cr_2O_7}} \times \frac{18.02\ g}{1\ \cancel{mol\ H_2O}}$$

$$= 0.858\ g$$

When the reaction is complete, the mass of Cr_2O_3 is 1.81 g, the mass of N_2 is 0.333 g, and the mass of H_2O is 0.858 g.

2. Textbook question 30:

A sample answer may be as follows: A lab technician from the city of Edmonton was very excited to relate aspects of her daily routines. Among many duties, she checks for fluoride, chloride, and magnesium in the water supply. Chloride is checked with a device called an amperometric titrater. This works by liberating a chemical and measuring a change in the charge on an ampmeter.

Chlorine is added to the water at a rate of 2 ppm (parts per million) in the form of chloroamine (NH_2Cl) – a molecular compound. A typical day in Edmonton sees a consumption of 350 000 000 L of water.

Fluoride is measured with a pH meter with an ion-specific electrode. Fluoride is added to Edmonton water at a rate of 1 ppm.

When chemicals are added to the water, ions such as magnesium and calcium are released and measured. Magnesium is measured with an ion chromatograph. Water passes through resin in a column. Magnesium reacts with a chemical and is slowed down, then is reflected as a peak on a graph.

Lime (CaO) is also added to soften the water and raise the pH level. The softening effect occurs by the calcium reacting with carbonates in the water to precipitate calcium carbonate. This removes the hardening effect of the calcium ions from the water.

Textbook question 32:

The education required to become an analytical chemist is as follows:

- Have a minimum four-year degree in chemistry or biochemistry.
- Have courses in inorganic and organic chemistry, mathematics (including calculus), physics, and the humanities.

The following is a list of possible work places:

- government labs
- water treatment plants
- forensic labs
- university labs
- food processing plants
- pharmaceutical labs

A sample interview may be as follows:

Q: What kind of training do you have?

A: I studied chemistry at the University of Calgary, and biochemistry at McMaster University in Hamilton.

Q: What kinds of things do you do on a routine basis?

A: Actually, I cut out the intestines of mice that have cancer, then I look for certain chemicals that might be found in the linings of the intestines.

Q: What kind of quantities are you looking for?

A: We work with extremely small amounts called micromoles. That is still a very large number of molecules, but in terms of mass, it's a very, very small amount.

Q: How much money do you make and how secure is your job?

A: When you work in a research lab, your job security depends on grants and on getting some kind of results. When you work in industry, your job security depends on the success of the company. Salaries as an analytical chemist can vary widely. A starting salary for a chemist working for a chemical manufacturing company might be around $30 000 per year. The salary would then increase depending on experience and extra education obtained.

Section 2: Activity 1

1. The two steps common to all stoichiometry problems are the following:

   - the balanced chemical equation
   - the mole ratio calculation

2. Since volume units are usually litres (L), and the mole unit is mol, the conversion will have mol on top and L on the bottom.

$$\not{L} \times \boxed{\frac{\text{mol}}{\not{L}}} = \text{mol}$$

3. The method applies to all chemical reactions.

4. The general steps for both solids and gases are the same.

5. You need moles of propane in order to find moles of oxygen through the mole ratio.

6. L/mol allows you to cancel moles and obtain litres, the required answer.

7. Textbook question 18:

$$2\,CH_3OH_{(l)} + 3\,O_{2(g)} \rightarrow 2\,CO_{2(g)} + 4\,H_2O_{(g)}$$

15 g	v = ?
32.05 g/mol	24.8 L/mol

$$n_{(CH_3OH)} = 15\ \cancel{g\ CH_3OH} \times \frac{1\ mol}{32.05\ \cancel{g\ CH_3OH}}$$

$$= 0.47\ mol$$

$$n_{(O_2)} = 0.47\ \cancel{mol\ CH_3OH} \times \frac{3\ mol\ O_2}{2\ \cancel{mol\ CH_3OH}}$$

$$= 0.70\ mol\ O_2$$

$$v_{(O_2)} = 0.70\ \cancel{mol\ O_2} \times \frac{24.8\ L}{1\ \cancel{mol\ O_2}}$$

$$= 17\ L$$

or, in one line, $v_{(O_2)} = 15\ \cancel{g\ CH_3OH} \times \frac{1\ \cancel{mol\ CH_3OH}}{32.05\ \cancel{g\ CH_3OH}} \times \frac{3\ \cancel{mol\ O_2}}{2\ \cancel{mol\ CH_3OH}} \times \frac{24.8\ L}{1\ \cancel{mol\ O_2}}$

$$= 17\ L$$

The volume of oxygen needed is 17 L.

Textbook question 19:

The only number given that is important to you is the 50 L O_2.

$C_3H_{8(g)}$	+	$5\ O_{2(g)}$	→	$3\ CO_{2(g)}$	+	$4\ H_2O_{(g)}$
$m = ?$		50 L				
44.11 g/mol		24.8 L/mol				

$$n_{(O_2)} = 50\ \cancel{L\ O_2} \times \frac{1\ mol}{24.8\ \cancel{L\ O_2}}$$

$$= 2.0\ mol$$

$$n_{(C_3H_8)} = 2.0\ \cancel{mol\ O_2} \times \frac{1\ mol\ C_3H_8}{5\ \cancel{mol\ O_2}}$$

$$= 0.40\ mol\ C_3H_8$$

$$m_{(C_3H_8)} = 0.40\ \cancel{mol\ C_3H_8} \times \frac{44.11\ g}{1\ \cancel{mol\ C_3H_8}}$$

$$= 18\ g$$

or, in one line, $m_{(C_3H_8)} = 50\ \cancel{L\ O_2} \times \frac{1\ \cancel{mol\ O_2}}{24.8\ \cancel{L\ O_2}} \times \frac{1\ \cancel{mol\ C_3H_8}}{5\ \cancel{mol\ O_2}} \times \frac{44.11\ g}{1\ \cancel{mol\ C_3H_8}}$

$$= 18\ g$$

The mass of propane that can be burned is 18 g.

8. $2\,SO_{2(g)} + O_{2(g)} \rightarrow 2\,SO_{3(g)}$

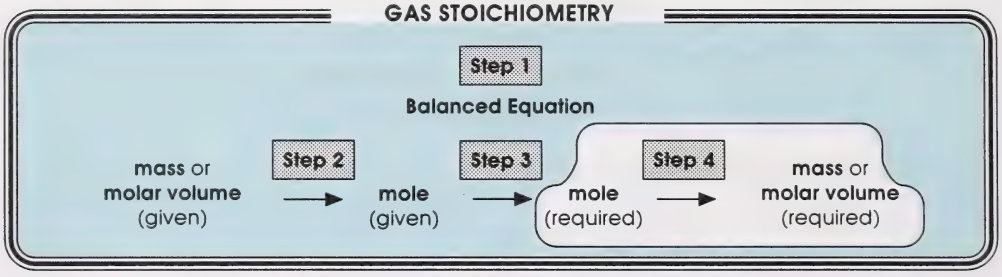
15.75 L	v = ?
22.4 L/mol	22.4 L/mol

$$n_{(SO_2)} = 15.75\ \cancel{L\ SO_2} \times \frac{1\ \cancel{mol\ SO_2}}{22.4\ \cancel{L\ SO_2}} \times \frac{1\ \cancel{mol\ O_2}}{2\ \cancel{mol\ SO_2}} \times \frac{22.4\ L}{1\ \cancel{mol\ O_2}}$$

$$= 7.88\ L$$

The amount of oxygen required is 7.88 L.

9. a.



b. The Ideal Gas Law could also be used in the first calculation step for finding moles of given (Step 2).

10. $2\,NaN_{3(s)} \rightarrow 2\,Na_{(s)} + 3\,N_{2(s)}$

58.0 g	v = ?
65.02 g/mol	20°C, 104.8 kPa

$$n_{(NaN_3)} = 58.0\ \cancel{g\ NaN_3} \times \frac{1\ mol\ NaN_3}{65.02\ \cancel{g\ NaN_3}}$$

$$= 0.892\ mol\ NaN_3$$

$$n_{(N_2)} = 0.892\ \cancel{mol\ NaN_3} \times \frac{3\ mol\ N_2}{2\ \cancel{mol\ NaN_3}}$$

$$= 1.34\ mol\ N_2$$

$$pv = nRT$$

$$v_{(N_2)} = \frac{nRT}{p}$$

$$v_{(N_2)} = \frac{1.34\ mol\ N_2 \times \frac{8.31\ kPa \cdot L}{mol \cdot K} \times 293\ K}{104.8\ kPa}$$

$$= 31.1\ L\ N_2$$

The volume of $N_{2(g)}$ produced is 31.1 L.

11. The evidence may be charted as follows:

Mass of Mg	0.054 g
Initial volume of gas	0.0 mL
Final volume of gas	54.8 mL
Temperature	21.3°C
Pressure	99.94 kPa

12. Because of the reactants involved, the gas produced must have been $H_{2(g)}$.

$$\begin{array}{r} \text{final volume } H_{2(g)} = 54.8 \text{ mL} \\ \text{initial volume } H_{2(g)} = \ \ 0.0 \text{ mL} \\ \hline \text{volume of } H_{2(g)} = 54.8 \text{ mL} \end{array}$$

$$\begin{array}{ll} Mg_{(s)} + 2\,HCl_{(aq)} \rightarrow H_{2(g)} + MgCl_{2(aq)} & \\ 0.054 \text{ g} & 54.8 \text{ mL} \\ 24.31 \text{ g/mol} & 21.3°C \\ & 99.94 \text{ kPa} \end{array}$$

$$n_{(Mg)} = 0.054 \text{ g Mg} \times \frac{1 \text{ mol Mg}}{24.31 \text{ g Mg}}$$

$$= 0.0022 \text{ mol Mg}$$

$$n_{(H_2)} = 0.0022 \text{ mol Mg} \times \frac{1 \text{ mol } H_2}{1 \text{ mol Mg}}$$

$$= 0.0022 \text{ mol } H_2$$

$$pv = nRT \rightarrow R = \frac{pv}{nT}$$

$$R = \frac{99.94 \text{ kPa}(0.0548 \text{ L})}{0.0022 \text{ mol}(294.2 \text{ K})}$$

$$= 8.5 \frac{\text{kPa} \cdot \text{L}}{\text{mol} \cdot \text{K}}$$

13. $$\% \text{ difference} = \frac{|\text{experimental value} - \text{predicted value}|}{\text{predicted value}} \times 100$$

$$= \frac{|8.5 - 8.31| \ \frac{\text{kPa} \cdot \text{L}}{\text{mol} \cdot \text{K}}}{8.31 \ \frac{\text{kPa} \cdot \text{L}}{\text{mol} \cdot \text{K}}} \times 100$$

$$= 2.3\%$$

Based on the percent difference of less than 10%, the experimental design and procedure would be judged to be sound or adequate.

Possible sources of error include the following:

- Some of the gas is lost when trying to collect the gas.
- There may be a few air bubbles in the cylinder to start with, making the gas volume greater than it should be.
- The water vapour that forms as the hydrogen gas bubbles into the cylinder is not taken into account. This would make the gas volume of H_2 appear to be more than it should.

14. The evidience may be charted as follows:

Mass of Mg	0.054 g
Initial volume of gas	0.0 mL
Final volume of gas	54.8 mL
Temperature	21.3°C
Pressure	99.94 kPa

15. Because of the reactants involved, the gas produced must have been $H_{2(g)}$.

$$\begin{aligned} \text{final volume } H_{2(g)} &= 54.8 \text{ mL} \\ \text{initial volume } H_{2(g)} &= 0.0 \text{ mL} \\ \hline \text{volume of } H_{2(g)} &= 54.8 \text{ mL} \end{aligned}$$

$$\begin{array}{ll} Mg_{(s)} + 2\,HCl_{(aq)} \rightarrow & H_{2(g)} + MgCl_{2(aq)} \\ 0.054 \text{ g} & 54.8 \text{ mL} \\ 24.31 \text{ g/mol} & 21.3°C \\ & 99.94 \text{ kPa} \end{array}$$

$$n_{(Mg)} = 0.054 \text{ g Mg} \times \frac{1 \text{ mol Mg}}{24.31 \text{ g Mg}}$$
$$= 0.0022 \text{ mol Mg}$$

$$n_{(H_2)} = 0.0022 \text{ mol Mg} \times \frac{1 \text{ mol } H_2}{1 \text{ mol Mg}}$$
$$= 0.0022 \text{ mol } H_2$$

$$pv = nRT \rightarrow R = \frac{pv}{nT}$$

$$R = \frac{99.94 \text{ kPa}(0.0548 \text{ L})}{0.0022 \text{ mol}(294.2 \text{ K})}$$

$$= 8.5 \frac{\text{kPa} \cdot \text{L}}{\text{mol} \cdot \text{K}}$$

16. $$\% \text{ difference} = \frac{|\text{experimental value} - \text{predicted value}|}{\text{predicted value}} \times 100$$

$$= \frac{|8.5 - 8.31| \frac{\text{kPa} \cdot \text{L}}{\text{mol} \cdot \text{K}}}{8.31 \frac{\text{kPa} \cdot \text{L}}{\text{mol} \cdot \text{K}}} \times 100$$

$$= 2.3\%$$

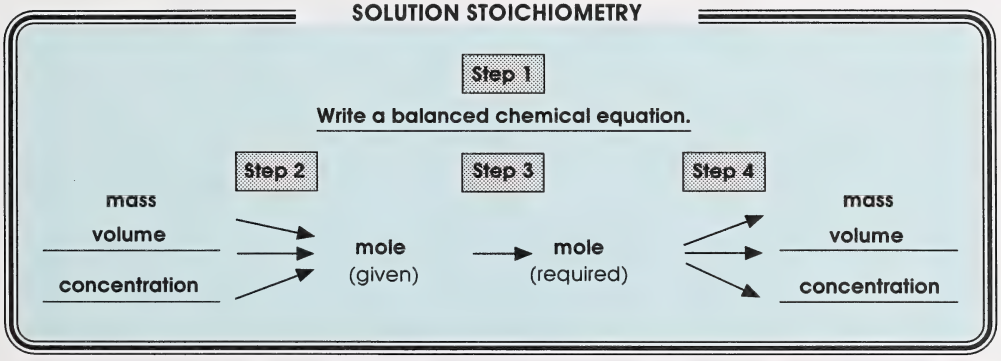
Based on the percent difference of less than 10%, the experimental design and procedure would be judged to be sound or adequate.

Possible sources of error include the following:

- Some of the gas is lost when trying to collect the gas.

- There may be a few air bubbles in the cylinder to start with, making the gas volume greater than it should be.
- The water vapour that forms as the hydrogen gas bubbles into the cylinder is not taken into account. This would make the gas volume of H_2 appear to be more than it should.

Section 2: Activity 2

1. Volume and molar concentration (sometimes mass) are involved in describing substances in solution.
2. Volume and molar concentration usually replace mass and molar mass in solution stoichiometry.
3. See the summary of general steps on page 181 of your textbook, *Nelson Chemistry*.
4. 

5. Textbook question 23:

$2\,NH_{3(aq)} + H_2SO_{4(aq)} \rightarrow (NH_4)_2SO_{4(aq)}$

24.4 mL 50.0 mL

2.20 mol/L C = ?

$$C_{(H_2SO_4)} = 24.4\ \cancel{mL} \times \frac{2.20\ \cancel{mol\ NH_3}}{L} \times \frac{1\ mol\ H_2SO_4}{2\ \cancel{mol\ NH_3}} \times \frac{1}{50.0\ \cancel{mL}}$$

$$= 0.537\ mol/L\ H_2SO_4$$

Note:

- The m (milli) was carried through and cancelled with the last term.
- The last term represents the reciprocal of dividing by 50.0 mL to get the concentration. It's just a way of fitting the calculation into the one-line method shown here.

The concentration is 0.537 mol/L H_2SO_4.

Textbook question 24:

$$3\ Ca(OH)_{2(aq)} + Al_2(SO_4)_{3(aq)} \rightarrow 3\ CaSO_{4(s)} + 2\ Al(OH)_{3(aq)}$$

v = ?	25.0 mL
0.0250 mol/L	0.125 mol/L

$$v_{(Ca(OH)_2)} = 25.0\ mL \times \frac{0.125\ \cancel{mol\ Al_2(SO_4)_3}}{\cancel{L}} \times \frac{3\ \cancel{mol\ Ca(OH)_2}}{1\ \cancel{mol\ Al_2(SO_4)_3}} \times \frac{1\ \cancel{L}}{0.0250\ \cancel{mol\ Ca(OH)_2}}$$

$$= 375\ mL$$

Note: The millilitre unit was carried through (not cancelled) to get a mL (millilitre) end volume.

The volume required is 375 mL $Ca(OH)_{2(aq)}$.

Textbook question 25:

$$3\ Na_2CO_{3(aq)} + 2\ FeCl_{3(aq)} \rightarrow 6\ NaCl_{(aq)} + Fe_2(CO_3)_{3(s)}$$

0.250 mol/L	75.0 mL
v = ?	0.200 mol/L

$$v_{(Na_2CO_3)} = 75.0\ mL \times \frac{0.200\ \cancel{mol\ FeCl_3}}{\cancel{L}} \times \frac{3\ \cancel{mol\ Na_2CO_3}}{2\ \cancel{mol\ FeCl_3}} \times \frac{1\ \cancel{L}}{0.250\ \cancel{mol\ Na_2CO_3}}$$

$$= 90.0\ mL$$

The minimum volume required is 90 mL $Na_2CO_{3(aq)}$.

6. $$3\ Na_2S_{(aq)} + 2\ Al(NO_3)_{3(aq)} \rightarrow 6\ NaNO_{3(aq)} + Al_2S_{3(s)}$$

20.0 mL	m = ?
0.210 mol/L	150.14 g/mol

$$m_{(Al_2S_3)} = 20.0\ m\cancel{L} \times \frac{0.210\ \cancel{mol\ Na_2S}}{\cancel{L}} \times \frac{1\ \cancel{mol\ Al_2S_3}}{3\ \cancel{mol\ Na_2S}} \times \frac{150.14\ g}{1\ \cancel{mol\ Al_2S_3}}$$

$$= 210\ mg$$

The prediction is that 210 mg of $Al_2S_{3(s)}$ will form.

7. The analysis may be outlined as follows:

mass of filter paper plus precipitate = 1.17 g
mass of filter paper = 0.97 g
---
mass of precipitate = 0.20 g (200mg)

8. A sample evaluation may be as follows:

The experimental design is adequate since it addresses the question as stated in the problem. By testing the filtrate, it is felt that the reaction had taken place fully, and that all the precipitate was measured. The excess aluminum nitrate in the filtrate indicated that the sodium sulphide was reacted completely. Therefore it was valid to make the prediction using the sodium sulphide measurements.

$$\% \text{ difference} = \frac{|200 \text{ mg} - 210 \text{ mg}|}{210 \text{ mg}} \times 100$$
$$= 4.76\%$$

The small percent difference supports the design and procedure as being adequate.

9. The implied lab test for silver nitrate is as follows:

Excess sodium sulphate in solution is added to a sample that may contain silver nitrate. A white precipitate forms.

10. The analysis may be outlined as follows:

mass of filter paper plus precipitate = 6.74 g
mass of filter paper = 1.27 g
---
mass of precipitate = 5.47 g

$$\underset{}{Na_2SO_{4(aq)}} + \underset{100 \text{ mL } (0.100 \text{ L})}{2\,AgNO_{3(aq)}} \rightarrow \underset{}{2\,NaNO_{3(aq)}} + \underset{5.47 \text{ g}}{Ag_2SO_{4(s)}}$$
$$\underset{C = ?}{} \qquad\qquad \underset{311.80 \text{ g/mol}}{}$$

$$C_{(AgNO_3)} = 5.47 \cancel{\text{ g } Ag_2SO_4} \times \frac{1 \cancel{\text{ mol } Ag_2SO_4}}{311.80 \cancel{\text{ g } Ag_2SO_4}} \times \frac{2 \text{ mol } AgNO_3}{1 \cancel{\text{ mol } Ag_2SO_4}} \times \frac{1}{0.100 \text{ L}}$$
$$= 0.351 \text{ mol/L } AgNO_3$$

11. Industry has decided that the concept has passed the test. The design is to test for the molar concentration of silver nitrate and industry is satisfied with the results.

12. $NaHCO_{3(aq)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)} + CO_{2(g)}$

50 mL(0.050 L)	excess	$v = ?$
0.200 mol/L		100 kPa
		35°C(308 K)

$$n_{(CO_2)} = 0.050 \cancel{L} \times \frac{0.200 \cancel{\text{mol NaHCO}_3}}{\cancel{L}} \times \frac{1 \text{ mol } CO_2}{1 \cancel{\text{mol NaHCO}_3}}$$

$$= 0.010 \text{ mol } CO_2$$

$$pv = nRT$$

$$v = \frac{nRT}{p}$$

$$v = \frac{0.010 \cancel{\text{mol}} \left(8.31 \frac{\cancel{\text{kPa}} \cdot \text{L}}{\cancel{\text{mol}} \cdot \cancel{\text{K}}}\right) 308 \cancel{\text{K}}}{100 \cancel{\text{kPa}}}$$

$$v = 0.26 \text{ L } CO_2$$

The volume of CO_2 is predicted to be 0.26 L or 2.6×10^2 mL.

Section 2: Activity 3

1. Sodium carbonate is one of the materials required in making soap and glass.

2. The LeBlanc process was used for producing soda ash, but the process required too much polluting coal and was expensive.

3. Following the principles of solubility, the Solvay process uses cold water to precipitate and separate the sodium hydrogen carbonate from the ammonium chloride solution.

4. Textbook question 33:

   a. $CaCO_{3(s)} + 2\,NaCl_{(aq)} \rightarrow Na_2CO_{3(aq)} + CaCl_{2(aq)}$

10.0 kL	$m = ?$
5.40 mol/L	105.99 g/mol

$$m_{(Na_2CO_3)} = 10.0 \cancel{\text{kL}} \times \frac{5.40 \cancel{\text{mol NaCl}}}{\cancel{L}} \times \frac{1 \cancel{\text{mol } Na_2CO_3}}{2 \cancel{\text{mol NaCl}}} \times \frac{105.99 \text{ g}}{1 \cancel{\text{mol } Na_2CO_3}}$$

$$= 2862 \text{ kg}$$

$$= 2.86 \times 10^3 \text{ kg (Round the answer to the correct number of significant digits (3).)}$$

The mass of soda ash that can be produced is 2.86×10^3 kg.

b. $CaCO_{3(s)} + 2\,NaCl_{(aq)} \rightarrow Na_2CO_{3(aq)} + CaCl_{2(aq)}$

500 kg $v = ?$

100.09 g/mol 5.40 mol/L

$$v_{(NaCl)} = 500\ \text{kg}\ \cancel{CaCO_3} \times \frac{1\ \cancel{\text{mol CaCO}_3}}{100.09\ \cancel{\text{g CaCO}_3}} \times \frac{2\ \cancel{\text{mol NaCl}}}{1\ \cancel{\text{mol CaCO}_3}} \times \frac{1\ \text{L}}{5.40\ \cancel{\text{mol NaCl}}}$$

$$= 1.85\ \text{kL}$$

Don't forget the k (kilo)!

The volume of brine solution required is 1.85 kL.

c. $CaCO_{3(s)} + 2\,NaCl_{(aq)} \rightarrow Na_2CO_{3(aq)} + CaCl_{2(aq)}$

350 L $v = ?$

2.00 mol/L 3.00 mol/L

$$v_{(CaCl_2)} = 350\ \cancel{\text{L}} \times \frac{2.00\ \cancel{\text{mol NaCl}}}{\cancel{\text{L}}} \times \frac{1\ \cancel{\text{mol CaCl}_2}}{2\ \cancel{\text{mol NaCl}}} \times \frac{1\ \text{L}}{3.00\ \cancel{\text{mol CaCl}_2}}$$

$$= 117\ \text{L}$$

The volume produced is 117 L $CaCl_2$.

5. Textbook question 21:

$2\,NaCl_{(l)} \rightarrow 2\,Na_{(s)} + Cl_{2(g)}$

100 kg $v = ?$ (SATP)

22.99 g/mol 24.8 L/mol

$$v_{(Cl_2)} = 100\ \text{kg}\ \cancel{\text{Na}} \times \frac{1\ \cancel{\text{mol Na}}}{22.99\ \cancel{\text{g Na}}} \times \frac{1\ \cancel{\text{mol Cl}_2}}{2\ \cancel{\text{mol Na}}} \times \frac{24.8\ \text{L}}{1\ \cancel{\text{mol Cl}_2}}$$

$$= 53.9\ \text{kL}$$

The volume of chlorine gas produced is 53.9 kL.

6. a. A percent difference greater than 10% would cause you to reject the experimental design as being inadequate.

b. Possible considerations in assessing experimental designs and the stoichiometric method would include the following:

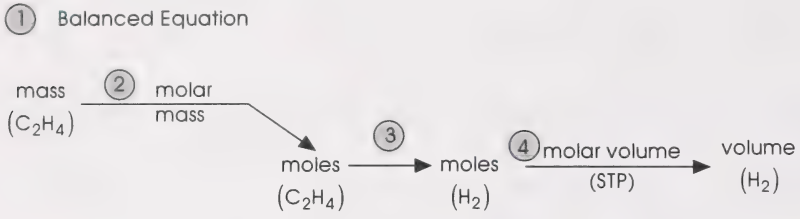
- There may have been a procedural error made in performing the experiment.
- There may be flaws in the equipment used in the experiment.
- The experimental design itself may be flawed.
- The stoichiometric method may need revising.

c. No, one exception would not be grounds for rejecting the method. However, if considerable scientific evidence were collected that challenged the stoichiometric method, revision or rejection could then be considered. This process would be the same in considering changes in scientific theories, as well.

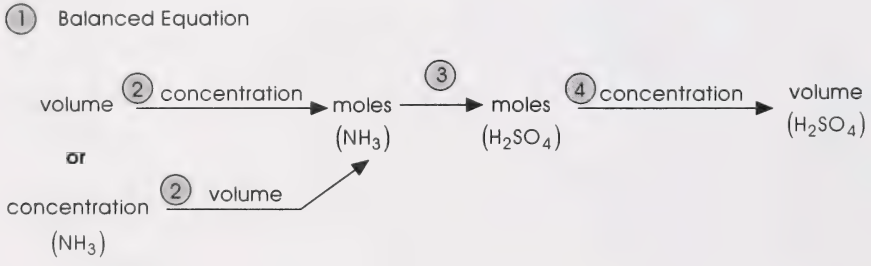
Section 2: Follow-up Activities

Extra Help

1. a. The pathway will look as follows:



b. The pathway will look as follows:



2. a. iii b. ii c. i

3. a.

$$v_{(H_2SO_4)} = 150\text{ mL NaOH} \times \frac{0.020\text{ mol NaOH}}{1\text{ L NaOH}} \times \frac{2\text{ mol NaOH}}{1\text{ mol } H_2SO_4} \times \frac{1\text{ L } H_2SO_4}{0.100\text{ mol } H_2SO_4}$$

(incorrect)

$$\boxed{150\text{ mL NaOH}} \times \boxed{\frac{0.020\text{ mol NaOH}}{1\text{ L}}} \times \boxed{\frac{\mathbf{1\text{ mol } H_2SO_4}}{\mathbf{2\text{ mol NaOH}}}} \times \boxed{\frac{1\text{ L } H_2SO_4}{0.100\text{ mol } H_2SO_4}}$$

The volume required is 15 mL H_2SO_4.

b.

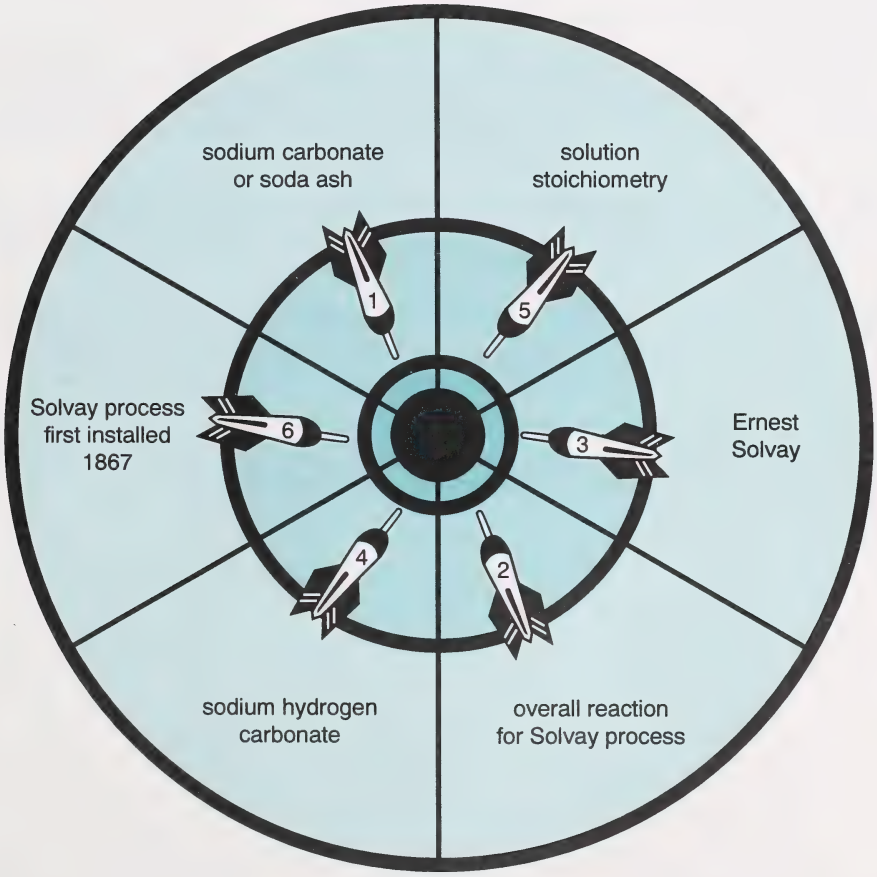
$$C_{(Pb(NO_3)_2)} = 48.5\ \text{mL}\ Na_2SO_4 \times \frac{1\ \text{L}}{0.050\ \text{mol}\ Na_2SO_4} \times \frac{1\ \text{mol}\ Pb(NO_3)_2}{1\ \text{mol}\ Na_2SO_4} \times 377\ \text{mol}\ Pb(NO_3)_2$$

incorrect

$$48.5\ \text{mL}\ Na_2SO_4 \times \frac{\mathbf{0.050\ mol\ Na_2SO_4}}{\mathbf{1\ L}} \times \frac{1\ \text{mol}\ Pb(NO_3)_2}{1\ \text{mol}\ Na_2SO_4} \times \frac{\mathbf{1}}{\mathbf{377\ mL\ Pb(NO_3)_2}}$$

The concentration of $Pb(NO_3)_2$ is 0.00643 mol/L.

4.



Enrichment

1. $C_6H_{12}O_{6(s)} + 6\,O_{2(g)} \rightarrow 6\,H_2O_{(l)} + 6\,CO_{2(g)}$

   1.00 g — 24.8 L/mol CO_2

   180.18 g/mol — v = ? (SATP)

$$v_{(CO_2)} = 1.00\ \cancel{g\ C_6H_{12}O_6} \times \frac{1\ \cancel{mol\ C_6H_{12}O_6}}{180.18\ \cancel{g\ C_6H_{12}O_6}} \times \frac{6\ \cancel{mol\ CO_2}}{1\ \cancel{mol\ C_6H_{12}O_6}} \times \frac{24.8\ L}{1\ \cancel{mol\ CO_2}}$$

$$= 0.826\ L$$

The volume of carbon dioxide released is 0.826 L.

2. At the end of the last century there was an ever-increasing demand for fertilizer to help make marginal farmland productive. The molecular form of nitrogen in the atmosphere is in a very large supply, but not in a useful form for fertilizer. In 1907-1909, Fritz Haber (1868-1934) developed a laboratory method for converting nitrogen into ammonia. He received a Nobel prize in 1918 for his efforts.

   An equally important contribution was made by Carl Bosch (1874-1940). He performed over 20 000 experiments with the Haber laboratory process that resulted in a way to use Haber's process on an industrial scale. He received a Nobel prize for his efforts in 1931.

3. Some metal hydrides form when metal crystals are treated with hydrogen gas to form nonstoichiometric compounds. These have formulas like $LaH_{2.76}$. When these compounds are heated, hydrogen is released. Because of this, these materials are being studied for possible use as portable fuel generators.

   The other interesting application of nonstoichiometric materials is in the formation of ceramic glazes where imperfections develop in the crystal lattice. The imperfections also allow the lead oxides to be electrical conductors. These nonstoichiometric compounds are used as electrodes in lead storage batteries.

NOTES

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